Chemistry & Chemical Reactivity 11th edition Kotz Solutions Manual

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Student Solutions Manual

Chemistry & Chemical Reactivity

ELEVENTH EDITION

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Solution and Answer Guide

Kotz Treichel Townsend Treichel, Chemistry and Chemical Reactivity 11e, 978-0-357-85140-1, Chapter 1: Basic Concepts of Chemistry

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Applying Chemical Principles

CO_2 in the Oceans

1.1.1. Name of CO_2: carbon dioxide

1.1.2. Symbols for metals mentioned in the article:

calcium, Ca; copper, Cu; manganese, Mn; iron, Fe

1.1.3. Most dense metal: Cu (8960 kg/m^3) Least dense metal: Ca (1550 kg/m^3)
Data taken from www.ptable.com

1.1.4. $CaCO_3$ (calcium carbonate) contains Ca (calcium), C (carbon), and O (oxygen).

Practicing Skills

Nature of Science

1.1. (a) Proposal that pressure increases with decreased volume—hypothesis

(b) Over time experiments indicate that pressure and volume are inversely proportional—law

(c) Proposal that more molecules colliding per given area results in increased pressure—theory

1.2. Categorize as hypothesis, theory, or law: Hypothesis--a tentative explanation or prediction in accord with current knowledge.

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Green Chemistry

1.3. Sustainable development means meeting today's needs while ensuring that future generations will be able to meet theirs.

1.4. Green chemistry refers to practices that reduce waste products during chemical processes, use materials wisely, use renewable materials, generate substances with the lowest possible toxicity, and conserve energy as well as materials.

1.5. Practices of Green Chemistry described:

- Synthetic methods to maximize incorporation of all materials
- Synthetic methods to generate substances with little or no toxicity
- Raw materials should be renewable
- Energy requirements recognized for environmental and economic impact minimized

1.6. Practices of Green Chemistry described:

- Raw materials (plant-based materials) renewable
- Energy saved
- Synthesis uses products with low or no toxicity (water and plant-based materials)
- Synthetic methods to generate substances with little or no toxicity

Matter: Elements and Atoms, Compounds and Molecules

1.7. The names of each of the elements:

(a)	N	nitrogen	(c)	Br	bromine	(e)	Li	lithium
(b)	Ca	calcium	(d)	I	iodine	(f)	Fe	iron – from the Latin *ferrium*

1.8. The names of each of the elements:

(a)	Cr	chromium	(c)	Mg	Magnesium – often confused with manganese	(e)	Ar	argon
(b)	Ni	nickel	(d)	Cl	chlorine	(f)	Ti	titanium

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1.9. The symbol for each of the elements:

(a)	strontium	Sr	(c)	cobalt	Co	(e)	selenium	Se
(b)	cadmium	Cd	(d)	mercury	Hg	(f)	bismuth	Bi

1.10. The symbol for each of the elements:

(a)	platinum	Pt	(c)	uranium	U	(e)	tungsten	W
(b)	gallium	Ga	(d)	thallium	Tl	(f)	xenon	Xe

1.11. In each of the pairs, decide which is an element and which is a compound:

[HINT: If the isolated symbol is on the periodic table, it's an element!]

(a) Na or NaCl—Sodium(Na) is an element, and sodium chloride(NaCl) is a compound.

(b) Sugar or carbon—Sugar, composed of C, H, and O, is a compound, and carbon(C) is an element.

(c) Gold or gold(III) chloride—Gold(Au) is an element, and gold(III) chloride ($AuCl_3$) is a compound.

1.12. In each of the pairs, decide which is an element and which is a compound:

[HINT: If the isolated symbol is on the periodic table, it's an element!]

(a) $Pt(NH_3)_2Cl_2$ is a compound; Pt is an element

(b) Copper is an element; copper(II) oxide is a compound

(c) Silicon is an element; SiO_2 is a compound

1.13. Masses of hydrogen and oxygen gases prepared from 27 g of water?

An 18 g sample of water contains 2 g of hydrogen gas and 16 g of oxygen gas. A 27 g sample will contain the same proportion of hydrogen and oxygen.

$$\frac{2 \text{ g hydrogen}}{18 \text{ g water}} = \frac{x}{27 \text{ g water}}$$

$$x = \frac{2 \times 27}{18} = 3$$

The amount of oxygen would be 27 – 3 or 24 g oxygen. We could have used the ratio of oxygen to water to solve for the amount of oxygen in 27 g water.

The Law of Constant Composition (or the Law of Definite Proportions) is used.

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1.14. 60. g of magnesium produces 100. g of magnesium oxide. A ratio will tell us the amount of oxide formed when 30. g of magnesium are used (An example of The Law of Constant Composition or the Law of Definite Proportions).

$$\frac{60.\text{ g magnesium}}{40.\text{ g oxygen}} = \frac{30.\text{ g magnesium}}{x}$$

$$x = \frac{30. \times 40.}{60.} = 20.\text{ g oxygen}$$

Physical and Chemical Properties

1.15. Determine if the property is a physical or chemical property for the following:

(a) color — a physical property

(b) transformed into rust — a chemical property

(c) explode — a chemical property

(d) density — a physical property

(e) melts — a physical property

(f) green — a physical property (as in (a))

Physical properties are those that can be observed or measured without changing the composition of the substance. Exploding or transforming into rust results in substances that are **different** from the original substances—and represent chemical properties.

1.16. Determine if the property is a physical or chemical property for the following:

[HINT: Physical changes do not alter the chemical composition of the substances whereas chemical changes produce new substances with different chemical compositions.]

(a) physical property—electrical conductivity does not affect the chemical state of copper

(b) physical property—viscosity is changed upon heating but it does not change the chemical state of olive oil (or other liquids)

(c) physical property—density is one of the basic physical properties of substances

(d) chemical property—fermentation is a chemical process that changes sugar into alcohol

(e) physical property—color of a substance is also a basic physical property

(f) chemical property—burning (combustion) converts gasoline to carbon dioxide and water substances that are different from gasoline

1.17. Determine if the change is a physical or chemical change for the following:

[HINT: Physical changes do not alter the chemical composition of the substances whereas chemical changes produce new substances with different chemical compositions.]

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(a) physical change—changes of state, from solid to liquid to gas and so forth, are physical changes

(b) chemical change—milk souring converts the milk sugars into acids

(c) chemical change—a chemical reaction occurs when baking soda and vinegar combine making a new product, carbon dioxide gas

(d) physical change—after sugar is dissolved in water we can evaporate the water then recover the sugar

(e) chemical change—color changes are frequently associated with chemical changes, in this case as the tomato ripens several chemical compounds change structure to produce a red, ripe tomato

(f) chemical change—colors produced by fireworks are the result of chemicals burning in air producing new chemical compounds and releasing energy in the form of light

1.18. Determine if the change is a physical or chemical change for the following:

[HINT: Physical changes do not alter the chemical composition of the substances whereas chemical changes produce new substances with different chemical compositions.]

(a) chemical change—color changes are frequently associated with chemical changes, in this case bleach reacts with the T-shirt dye changing its chemical structure

(b) physical change— changes of state, from solid to liquid to gas and so forth, are physical changes, in this case your exhaled breath is initially a gas that converts to a liquid in the cold air.

(c) chemical change—when one chemical compound, carbon dioxide, is converts to another chemical compound, sugar, a chemical reaction occurs

(d) physical change— changes of state, from solid to liquid to gas and so forth, are physical changes, in this case melting butter changes from a solid to a liquid

1.19. Which are physical properties or chemical properties:

[HINT: While a substance's color is a physical property a change in color is almost always a chemical change.]

(a) color and physical state are physical properties, reactivity is a chemical property

chlorine's color, yellow-green, and physical state, gas, are physical properties

chlorine's reaction with sodium is a chemical property

(b) color and physical state are physical properties, reactivity is a chemical property

sodium bicarbonate's color, white, and physical state, solid, are physical properties

sodium bicarbonate's reaction with an acid is a chemical property

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1.20. Which are physical properties or chemical properties:

[HINT: While a substance's color is a physical property a change in color is almost always a chemical change.]

(a) color, physical state, and density are physical properties, reactivity is a chemical property

copper(II) sulfide's color, black, density, 2.71 g/cm^3, and physical state, solid, are physical properties

copper(II) sulfide's reaction with an acid is a chemical property

(b) color and physical state are physical properties, reactivity is a chemical property

magnesium's color, silver, and physical state, metal, are physical properties

magnesium's reaction with oxygen is a chemical property

Energy

1.21. To move the lever, one uses mechanical energy. The energy resulting is manifest in electrical energy (which produces light); thermal (radiant) energy would be released as the bulb in the flashlight glows.

1.22. Mechanical energy propels the car, electrical energy recharges the batteries, (thermal) radiant energy is released as the sun shines on the solar panels.

1.23. Which represents potential energy and which represents kinetic energy:

(a) acoustic energy represents matter in motion (molecules in air are moving) — kinetic

(b) thermal energy represents matter in motion—kinetic

(c) gravitational energy represents the attraction of the earth for an object—and therefore energy due to position—potential

(d) chemical energy represents the energy stored in fuels—potential

(e) electrostatic energy represents the energy of separated charges—and therefore potential energy.

1.24. Kinetic to Potential or vice versa:

(a) Potential → kinetic as water falls

(b) Potential → kinetic as natural gas burns, air inside house is heated

(c) Kinetic → potential as electrons are stored during battery charge

(d) Potential → kinetic the potential energy of the chemicals is converted into kinetic energy of the moving piston

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1.25. Since 1500 J of energy is lost by the metal, the water must gain 1500 J of energy, as dictated by the Law of Conservation of Energy.

1.26. The energy lost by the falling book is gained by the floor (which typically doesn't move owing to a larger mass). Some of the energy is gained by surrounding air molecules in the form of sound.

General Questions

1.27. For the gemstone turquoise:

(a) Qualitative: blue-green color, solid state Quantitative: density; mass

(b) Extensive: Mass Intensive: Density; Color; Physical state

(c) Volume: $2.5\ \text{g}\left(\frac{1\ \text{cm}^3}{2.65\ \text{g}}\right) = 0.94\ \text{cm}^3$

1.28. Qualitative vs Quantitative observations; Extensive vs Intensive observations:

(a) Qualitative: shiny golden metallic appearance, crystals in form of perfect cubes

Quantitative: length of 0.40 cm on a side, mass of 0.32 g

(b) Extensive: Mass and length; Intensive: color, luster, and crystalline form

(c) Density = Mass/Volume = 0.32 g/(0.40 cm)^3 = 0.064 g/0.064 cm^3 = 5.0 g/cm^3 = 5.0 g/mL

1.29. Of the observations below, those which identify chemical properties:

[Chemical properties, in general, are those observed during a chemical change—as opposed to during a physical change.]

(a) Vinegar and sodium bicarbonate reaction –Chemical

(b) Sugar soluble in water--Physical

(c) Water boils at 100°C--Physical

(d) Ultraviolet light converts O_3 to O_2--Chemical

(e) Ice is less dense than water—Physical

1.30. Of the observations below, those which identify chemical properties:

[Chemical properties, in general, are those observed during a chemical change—as opposed to during a physical change.]

(a) Iron rusting—a chemical property as iron reacts with oxygen

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(b) Sodium metal reacts—a chemical property as sodium metal and water react

(c) Octane combustion—a chemical property as C_8H_{18} forms CO_2 and H_2O

(d) Chlorine is a yellow-green gas—a physical property (observable without a chemical reaction)

(e) Ice melting from heat—a physical property (observable without a chemical reaction)

1.31. Regarding fluorite:

(a) The symbols for the elements in fluorite: Ca (calcium) and F (fluorine);

(b) Shape of the crystals: cubic
Arrangement of ions in the crystal indicates that the fluoride ions are arranged around the calcium ions in the lattice in such a way as to form a cubic crystal.

1.32. Regarding azurite:

(a) Symbols of the elements: Copper, Cu; Carbon, C; Oxygen, O

(b) Oxygen is a gas, while copper, carbon, and azurite are solids at room temperature. Oxygen is colorless, while copper has a reddish color and carbon is gray/black. The gemstone is a bluish color clearly different from its component elements.

1.33. A solution is a mixture, so the components can be separated using a physical technique. If one heats the NaCl solution to dryness, evaporating all the water, the NaCl solid remains behind. Hence the physical property of boiling points is useful in this separation.

1.34. The non-uniform appearance of the mixture indicates that samples taken from different regions of the mixture would be different—a characteristic of a heterogeneous mixture. Recalling that iron is attracted to a magnetic field while sand is generally not attracted suggests that passing a magnet through the mixture would separate the sand and iron.

1.35. Identify physical or chemical changes:

(a) As there is no change in the composition of the carbon dioxide in the sublimation process, this represents a physical change.

(b) A change in density as a function of temperature does not reflect a change in the composition of the substance (mercury), so this phenomenon represents a physical change.

(c) The combustion of methane represents a change in the substance present as methane is converted to the oxides of hydrogen and carbon that we call water and carbon dioxide—a chemical change.

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(d) Dissolving NaCl in water represents a physical change as the solid NaCl ion pairs are separated by the solvent, water. This same phenomenon, the separation of ions, also occurs during melting.

1.36. Identify physical or chemical changes:

(a) The desalination of sea water represents a physical change—as the salts and solvent (water) are separated.

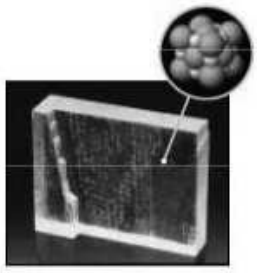
(b) The formation of SO_2 as sulfur-containing coal is burned represents a combination of sulfur and oxygen—a chemical change.

(c) The tarnishing of silver represents a chemical change as silver compounds form on the exterior of the silver object.

(d) Iron is heated to red heat. Changing the temperature of an object is a physical change.

1.37. A segment of Figure 1.2 is shown here:

The macroscopic view is the large crystal in the lower left of the figure, and the particulate view is the representation in the upper right. If one imagines reproducing the particulate (sometimes called submicroscopic) in all three dimensions—imagine a molecular duplicating machine—the macroscopic view results.



1.38. The orange solid and liquid in the bowl (top right) and the orange liquid and gas in the round-bottom flask (bottom right) represent the macroscopic views. The spheres (left column) represent the particulate view.

The particulate view of the solid displays molecules of bromine tightly packed to produce the solid.

The liquid view displays molecules of Br_2 that are still close together but not arranged in an orderly fashion and can move past one another.

The gas view displays molecules of Br_2, with the molecules being farther apart than in the liquid view.

1.39. A substance will float in any liquid whose density is greater than its own, and sink in any liquid whose density is less. The piece of plastic soda bottle ($d = 1.37\ g/cm^3$) will float in liquid CCl_4 and the piece of aluminum ($d = 2.70\ g/cm^3$) will sink in the liquid CCl_4.

1.40. Liquids: mercury and water; Solid: copper
Of the substances shown, mercury is most dense and water is least dense.

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1.41. Categorize each as an element, a compound, or a mixture:

(a) Iron pyrite (fool's gold)—is a compound composed of iron and sulfur

(b) Carbonated mineral water is a mixture. It certainly contains the compound water AND carbon dioxide. The term "mineral" implies that other dissolved materials are present.

(c) Molybdenum—an element

(d) Sucrose—a compound, with formula $C_{12}H_{22}O_{11}$

1.42. Categorize each as an element, a compound, or a mixture:

(a) Sea water is a mixture of water with many dissolved salts and gases.

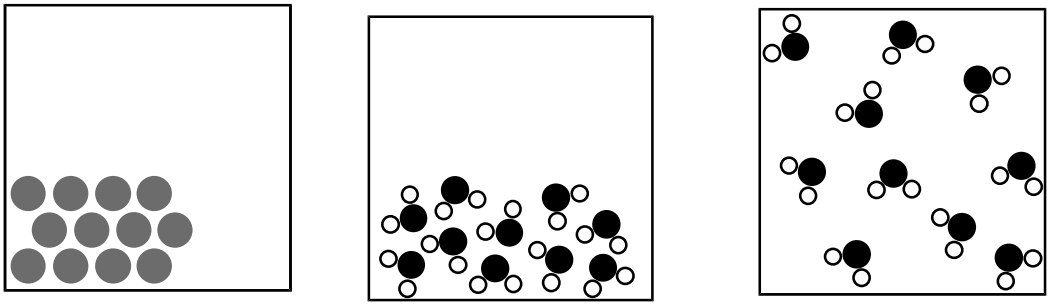
(b) Sodium chloride is a compound of sodium and chlorine.

(c) Bronze is a mixture of copper and tin.

(d) Gold (24-carat) is elemental gold refined to be at least 99.9% pure gold.

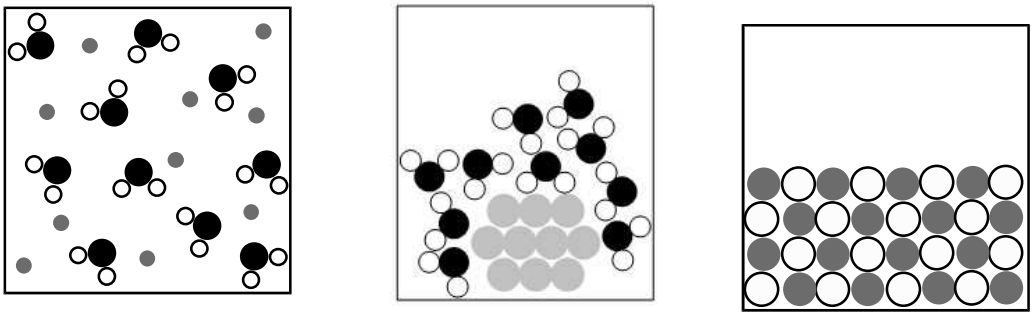
1.43. Indicate the relative arrangements of the particles in each of the following:

(a) solid iron (b) liquid water (c) gaseous water



1.44. Indicate the relative arrangements of the particles in each of the following:

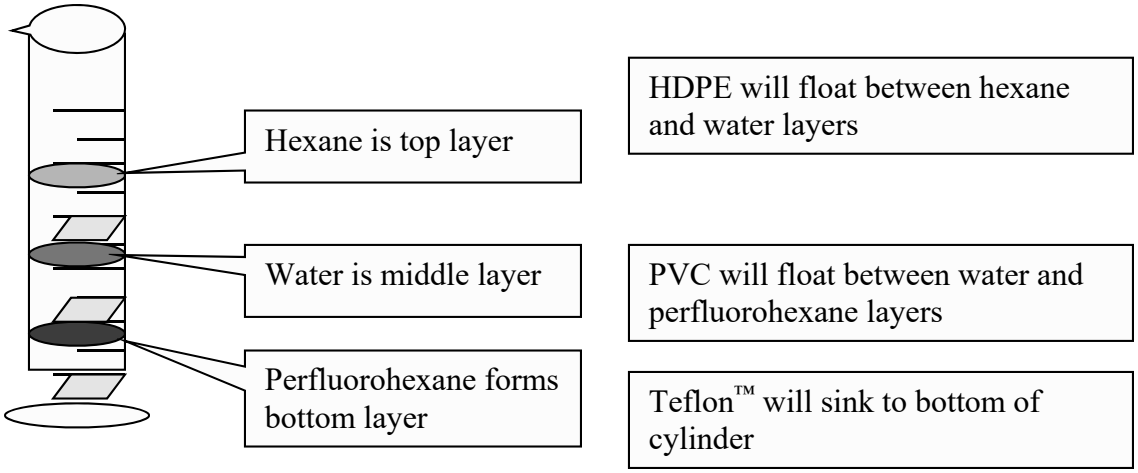
(a) $H_2O(g)$ & He(g) (b) $H_2O(\ell)$ & Al(s) (c) Cu(s) & Zn(s)



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1.45. When the three liquids are placed into the graduated cylinder, they will "assemble" in layers with increasingly smaller densities (from the bottom to the top) in the cylinder.



[Shadings are added to the top of each layer to provide clarity only (NOT to indicate the color of the liquids). Similarly the parallelogram symbols indicating the plastic samples are shaded—only to provide clarity in locating them—and not to imply any specific colors.]

1.46. There are several properties that could be used. One is the melting point. Sugar melts around 160-186 °C while table salt melts about 800 °C.

1.47. HDPE with a density of 0.97 g/cm^3 will float in any liquid whose density is greater than 0.97 g/ cm^3 and sink in any liquid whose density is less than that of HDPE. Of the liquids listed, HDPE should float in ethylene glycol, water, acetic acid, and glycerol.

1.48. Measure the density and melting point of the silvery metal and compare to published data points for silver.

1.49. Water, a large component of milk, expands as it is converted to the solid state—as the density of solid water is less than that of liquid water. The expanded solid escapes via the avenue of least resistance—the cap.

1.50. First, weigh the object. Then immerse it in a liquid (e.g. in a graduated cylinder) in which it sinks. If you measure the volume of liquid before and after you immerse the metal, the difference in volume is the volume of the object. Calculate the density of the object by dividing the mass by its volume.

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1.51. If one excretes too much sugar, the concentration of the sugar "solution" left in the body would decrease, resulting in urine with a higher density. If one excretes too much water, the concentration of the sugar "solution" in the body would increase, with a concomitant decrease in density of the urine.

1.52. To determine the identity, use the physical properties of water: measure the density of the unknown liquid. Freeze and boil the liquid. If the density of the liquid is about 1g/cm^3, if the liquid freezes about 0 ˚C and boils about 100 ˚C, the liquid is probably water. To test for the presence of salt, use a conductivity device. Water containing dissolved salts will conduct an electric current while pure water will not.

1.53. For the reaction of elemental potassium reacting with water:

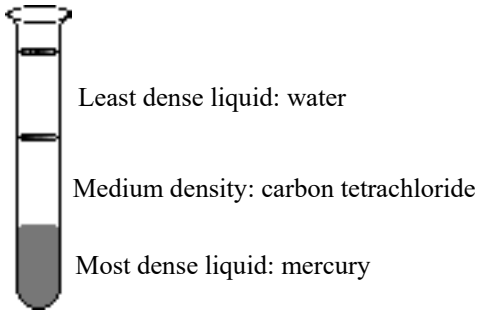
(a) States of matter involved: **Solid** potassium reacts with **liquid** water to produce **gaseous** hydrogen and **aqueous** potassium hydroxide solution (a homogenous mixture).

(b) The observed change is **chemical**. The products (hydrogen and potassium hydroxide) are quite different from elemental potassium and water. Litmus paper would also provide the information that while the original water was neither acidic nor basic, the solution produced would be basic. (The color of red litmus paper would change to blue.)

(c) The **reactants**: potassium and water
The **products**: hydrogen, potassium hydroxide solution, heat, and light

(d) Potassium reacts **vigorously** with water. Potassium is less dense than water, and floats atop the surface of the water. The reaction produces enough heat to ignite the hydrogen gas evolved. The flame observed is typically violet-purple in color. The potassium hydroxide formed is soluble in water (and therefore not visible).

1.54. The three liquids with the least dense liquid at the top and most dense liquid at bottom:



1.55. Since gases rise to an area with a similar density as their own, balloons with hydrogen and methane—with densities less than that of dry air--will float (and, if untethered, float away), while the balloons containing chlorine and propane will "sink"—to the lowest nearby surface.

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1.56. You cannot say that the metal is definitely copper based on the fact that it conducts electricity. Use the physical properties of copper. One could measure the density of the metal and compare it to the published density of copper. Other options are to determine the metal's boiling and melting points and then compare them to the published values.

1.57. The dissolution of iodine in ethanol (to make a solution) is a **physical** change, with iodine being the solute and ethanol the solvent.

1.58. For the mixture:

(a) Density of the mixture:

Calculate the mass of each substance (multiply density by volume):

Mass of $CHCl_3$: 10.0 mL · 1.492 g/mL = 14.92 g

Mass of $CHBr_3$: 5.0 mL · 2.890 g/mL = 14.45 g

The mass of the solution is: 14.92 g + 14.45 g or 29.37 g (29 to 2 significant figures) and a volume of 15.0 mL

The density of the mixture is then: 29 g/15.0 mL = 2.0 g/mL (to 2 significant figures).

(b) Density of yellow crystal:

Mass of solution = 2.07 g/mL · 20.0 mL = 41.4 g

Mass of solution = mass $CHCl_3$ + mass $CHBr_3$

Let x = volume $CHCl_3$ and (20.0 mL − x) = volume $CHBr_3$

Then, mass of solution = (1.492 g/mL)x + (2.890 g/mL)(20.0 mL – x) = 41.4 g

x = 11.7 mL $CHCl_3$ and (20.0 – x) = 8.3 mL $CHBr_3$

So to obtain 20.0 mL of solution with d = 2.07 g/cm^3, mix 11.7 mL $CHCl_3$ and 8.3 mL $CHBr_3$.

1.59. For the wedding band in question:

(a) Gold- Au; Copper- Cu; Silver- Ag

(b) The reported density of Osmium (according to www.ptable.com) is 22.59 g/cm^3, making it the densest element.

(c) As the band is 18-carat (or 75% gold), we can multiply the mass of the ring by 75% written as a fraction.

$$5.58 \text{ g ring}\left(\frac{75 \text{ g Au}}{100 \text{ g ring}}\right) = 4.185 \text{ g Au or } 4.2 \text{ g Au to 2 significant figures}$$

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(d) Mass of lost gold:

$$112\times10^{6}\text{ rings}\left(\frac{6.15\ \times10^{-3}\text{ g Au lost}}{1\text{ ring}}\right)=6.89\times10^{5}\text{g Au lost(3 significant figures)}$$

Value of lost gold:

$$6.89\times10^{5}\text{g Au lost}\left(\frac{1\text{troy oz}}{31.1\text{ g Au}}\right)\left(\frac{\$1815}{1\text{troy oz}}\right)=\$4.02\times10^{7}\text{or \$40.2 million (3 significant figures)}$$

Note the assumption in the first calculation that each married couple would have 2 rings.

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Solution and Answer Guide

Kotz Treichel Townsend Treichel, Chemistry and Chemical Reactivity 11e, 978-0-357-85140-1, Chapter 1R: Let's Review

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Applying Chemical Principles

Out of Gas!

1R.1. Fuel density in kg/L: $\frac{1.77\ \text{lb}}{1\ \text{L}} \cdot \frac{0.4536\ \text{kg}}{1\ \text{lb}} = 0.803\ \text{kg/L}$

1R.2. Mass and volume of fuel that should have been loaded:

Mass of fuel already in tank: $7682\ \text{L} \times \frac{0.803\ \text{kg}}{\text{L}} = 6170\ \text{kg}$

Mass of fuel needed: $22{,}300\ \text{kg} - 6170\ \text{kg} = 16{,}100\ \text{kg}$

Volume of fuel needed: $16{,}100\ \text{kg} \times \frac{1\ \text{L}}{0.803\ \text{kg}} = 20{,}100\ \text{L}$

Ties in Swimming and Significant Figures

1R.1. Distance traveled in 0.001 s (at world record rate of 20.91 s for 50-m):

$$\frac{50\ \text{m}}{20.91\ \text{s}} \cdot \frac{0.001\ \text{s}}{1} \cdot \frac{1000\ \text{mm}}{1\ \text{m}} = 2.4\ \text{mm (2sf)}$$

1R.2. Time to travel 3.0 cm at world record rate:

$$\frac{20.91\ \text{s}}{50\ \text{m}} \cdot \frac{1\ \text{m}}{100\ \text{cm}} \cdot \frac{3.0\ \text{cm}}{1} = 0.0125\ \text{s or } 0.013\ \text{s (2sf)}$$

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1R.3. Percent error if lane is 3 cm longer than 50.00 m

Actual length of lane is 50.03 m (3 cm + 50.00 m converted to units of m)

Error = 50.03 m – 50.00 m = 0.03 m

$$\% \text{ error} = \frac{\text{error}}{\text{accepted value}} \times 100\% = \frac{0.03 \text{ m}}{50.00} \times 100\% = 0.06\%$$

Practicing Skills

Temperature Scales

1R.1. Express 25 ˚C in kelvins:

$$T(\text{K}) = (25\ ^\circ\text{C} + 273.15\ ^\circ\text{C}) \frac{1\ \text{K}}{1\ ^\circ\text{C}} = 298\ \text{K}$$

1R.2. Express 5.5×10^3 °C in kelvins: $(5.5 \times 10^3\ ^\circ\text{C} + 273.15\ ^\circ\text{C}) \frac{1\ \text{K}}{1\ ^\circ\text{C}} = 5.8 \times 10^3\ \text{K}$

1R.3. Make the following temperature conversions:

˚C	K
(a) 16	16 + 273.15 = 289
(b) 370 − 273 or 97	370
(c) 40	40 + 273.15 = 310

Note no decimal point after 40

1R.4. Make the following temperature conversions:

˚C	K
(a) 77 − 273.15 = –196	77
(b) 63	63 + 273.15 = 336
(c) 1450 – 273.15 =1177	1450

Length, Volume, Mass, and Density

1R.5. The distance of a marathon (42.195 km) in meters; in miles:

$$\frac{42.195 \text{ km}}{1} \cdot \frac{1000 \text{ m}}{1 \text{ km}} = 42195 \text{ m} \qquad \frac{42.195 \text{ km}}{1} \cdot \frac{0.62137 \text{ miles}}{1 \text{ km}} = 26.219 \text{ miles}$$

The factor (0.62137 mi/km) is found inside the back cover of the text.

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1R.6. Length of a 19 cm pencil in mm; in m:

$$\frac{19\ \text{cm}}{1} \cdot \frac{10\ \text{mm}}{1\ \text{cm}} = 190\ \text{mm} \qquad \frac{19\ \text{cm}}{1} \cdot \frac{1\ \text{m}}{100\ \text{cm}} = 0.19\ \text{m}$$

1R.7. Express the area of a 2.5 cm × 2.1 cm stamp in cm^2 ; in m^2 :

$$2.5\ \text{cm} \cdot 2.1\ \text{cm} = 5.3\ \text{cm}^2$$

$$\frac{5.3\ \text{cm}^2}{1} \cdot \left(\frac{1\ \text{m}}{100\ \text{cm}}\right)^2 = 5.3 \times 10^{-4}\ \text{m}^2$$

1R.8. Surface area of a cat food can's lid in cm^2 and m^2:

The radius of the lid is half the diameter.

$$\text{Area} = \pi r^2 = \pi\left(\frac{8.3\ \text{cm}}{2}\right)^2 = 54\ \text{cm}^2$$

$$8.3\ \text{cm} \cdot \frac{1\ \text{m}}{100\ \text{cm}} = 0.083\ \text{m}$$

$$\text{Area} = \pi r^2 = \pi\left(\frac{0.083\ \text{m}}{2}\right)^2 = 0.0054\ \text{m}^2$$

1R.9. Express volume of 250. mL beaker in cm^3; in liters (L); in m^3; in dm^3:

$$\frac{250.\ \text{mL}}{1\ \text{beaker}} \cdot \frac{1\ \text{cm}^3}{1\ \text{mL}} = \frac{250.\ \text{cm}^3}{1\ \text{beaker}}$$

$$\frac{250.\ \text{cm}^3}{1\ \text{beaker}} \cdot \frac{1\ \text{L}}{1000\ \text{cm}^3} = \frac{0.250\ \text{L}}{1\ \text{beaker}}$$

$$\frac{250.\ \text{cm}^3}{1\ \text{beaker}} \cdot \frac{1\ \text{m}^3}{1\times10^6\ \text{cm}^3} = \frac{2.50 \times 10^{-4}\ \text{m}^3}{1\ \text{beaker}}$$

$$\frac{250.\ \text{cm}^3}{1\ \text{beaker}} \cdot \frac{1\ \text{L}}{1000\ \text{cm}^3} \cdot \frac{1\ \text{dm}^3}{1\ \text{L}} = \frac{0.250\ \text{dm}^3}{1\ \text{beaker}}$$

1R.10. Volume of bottle in mL; in cm^3 ; in dm^3:

$$1.5\ \text{L} \cdot \frac{10^3\ \text{mL}}{1\ \text{L}} = 1.5 \times 10^3\ \text{mL}$$

$$1.5\ \text{L} \cdot \frac{10^3\ \text{mL}}{1\ \text{L}} \cdot \frac{1\ \text{cm}^3}{1\ \text{mL}} = 1.5 \times 10^3\ \text{cm}^3$$

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$$1.5\ \text{L} \cdot \frac{10^3\ \text{mL}}{1\ \text{L}} \cdot \frac{1\ \text{cm}^3}{1\ \text{mL}} \cdot \left(\frac{1\ \text{dm}}{10\ \text{cm}}\right)^3 = 1.5\ \text{dm}^3$$

1R.11. Convert book's mass of 2.52 kg into grams:

$$\frac{2.52\ \text{kg}}{1\ \text{book}} \cdot \frac{1 \times 10^3\text{g}}{1\ \text{kg}} = \frac{2.52 \times 10^3\ \text{g}}{1\ \text{book}}$$

1R.12. Mass of 2.265 g in kg; in mg:

$$2.265\ \text{g} \cdot \frac{1\ \text{kg}}{1 \times 10^3\text{g}} = 2.265 \times 10^{-3}\text{kg}$$

$$2.265\ \text{g} \cdot \frac{1 \times 10^3\ \text{mg}}{1\ \text{g}} = 2.265 \times 10^3\ \text{mg}$$

1R.13. What mass of ethylene glycol (in grams) possesses a volume of 500. mL of the liquid?

$$\frac{500.\ \text{mL}}{1} \cdot \frac{1\ \text{cm}^3}{1\ \text{mL}} \cdot \frac{1.11\ \text{g}}{\text{cm}^3} = 555\ \text{g}$$

1R.14. Mass of 100. mL of acetone:

Remember that 1 mL = 1 cm^3

$$100.\ \text{cm}^3\left(\frac{0.7845\ \text{g}}{1\ \text{cm}^3}\right) = 78.5\ \text{g}$$

1R.15. Volume of 2.365 g of silver: $2.365\ \text{g} \cdot \frac{1\ \text{cm}^3}{10.5\ \text{g}} = 0.225\ \text{cm}^3$

1R.16. Volume of 50.0 g of lead: $50.0\ \text{g} \cdot \frac{1\text{cm}^3}{11.35\ \text{g}} = 4.41\ \text{cm}^3$

1R.17. To determine the density, given the data, one must first convert each length to units of cm and then calculate the volume:

$$\left[\frac{1.05\ \text{mm}}{1} \cdot \frac{1\ \text{cm}}{10\ \text{mm}}\right] \cdot \frac{2.35\ \text{cm}}{1} \cdot \frac{1.34\ \text{cm}}{1} = 0.3306\ \text{cm}^3\ (0.331 \text{ to 3sf})$$

Density is Mass/Volume or $\frac{2.361\ \text{g}}{0.3306\ \text{cm}^3} = 7.14\ \frac{\text{g}}{\text{cm}^3}$.

Given the selection of metals, the identity of the metal is **zinc**.

1R.18. Larger volume, 600 g of water or 600 g of lead:

$$600\ \text{g}\ H_2O \cdot \frac{1\ \text{cm}^3}{0.995\ \text{g}} = 600\ \text{cm}^3 \qquad 600\ \text{g lead} \cdot \frac{1\ \text{cm}^3}{11.35\ \text{g}} = 50\ \text{cm}^3$$

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Energy Units

1R.19. Express the energy of a 1200 Calories/day diet in joules:

$$\frac{1200\ \text{Cal}}{1\ \text{day}} \cdot \frac{1000\ \text{calorie}}{1\ \text{Cal}} \cdot \frac{4.184\ \text{J}}{1\ \text{cal}} = 5.0 \times 10^6\ \text{Joules/day}$$

1R.20. Express 1670 kJ in dietary Calories: $1670\ \text{kJ} \cdot \frac{10^3\text{J}}{1\ \text{kJ}} \cdot \frac{1\ \text{cal}}{4.184\ \text{J}} \cdot \frac{1\ \text{Cal}}{10^3\text{cal}} = 399\ \text{Cal}$

1R.21. Compare 170 kcal/serving and 280 kJ/serving.

$$\frac{170\ \text{kcal}}{1\ \text{serving}} \cdot \frac{1000\ \text{calorie}}{1\ \text{kcal}} \cdot \frac{4.184\ \text{J}}{1\ \text{cal}} \cdot \frac{1\ \text{kJ}}{1000\ \text{J}} = 710\ \text{kJoules/serving}$$

So, 170 kcal/serving has a greater energy content.

1R.22. Greater energy per mL, berry juice or soft drink:

$$\text{Soft drink: } 130\ \text{Cal} \cdot \frac{10^3\text{cal}}{1\ \text{Cal}} \cdot \frac{4.184\ \text{J}}{1\ \text{cal}} = 5.4 \times 10^5\text{J} \text{ or } 540\ \text{kJ}$$

With the berry juice providing 630 kJ, the JUICE provides more energy.

$$\frac{540\ \text{kJ}}{335\ \text{mL}} = \frac{1.6\ \text{kJ}}{1\ \text{mL}} \qquad \frac{630\ \text{kJ}}{295\ \text{mL}} = \frac{2.1\ \text{kJ}}{1\ \text{mL}}$$

The juice also provides more energy per milliliter.

Accuracy, Precision, Error, and Standard Deviation

1R.23. Using the data provided, the averages and their deviations are as follows:

Data point	Method A	deviation	Method B	deviation
1	2.2	–0.2	2.703	0.000
2	2.3	–0.1	2.701	–0.002
3	2.7	0.3	2.705	0.002
4	2.4	0.0	2.703	0.000
Averages:	2.4		2.703	

Note that the deviations for both methods are calculated by first determining the average of the four data points, and then subtracting the individual data points from the average.

(a) The average density for method A is 2.4 g/cm^3 while the average density for method B is 2.703 g/cm^3

(b) The percent error for each method:

Error = experimental value - accepted value

From Method A error = (2.4 – 2.702) = –0.3 g/cm^3

From Method B error = (2.703 – 2.702) = 0.001 g/cm^3

and the percent error is then:

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$$(\text{Method A}) = \frac{-0.3}{2.702} \cdot 100\% = -10\% \text{ (1 sf)}$$

$$(\text{Method B}) = \frac{0.001}{2.702} \cdot 100\% = 0.04\% \text{ (1 sf)}$$

(c) The standard deviation for each method:

Method A: $\sqrt{\frac{(-0.2)^2 + (-0.1)^2 + (0.3)^2 + (0.0)^2}{3}} = \sqrt{\frac{0.14}{3}} = 0.2 \text{ g/cm}^3 \text{ (1 sf)}$

Method B:

$$\sqrt{\frac{(0.000)^2 + (-0.002)^2 + (0.002)^2 + (0.000)^2}{3}} = \sqrt{\frac{0.000008}{3}} = 0.002 \text{ g/cm}^3 \text{ (1 sf)}$$

(d) Method B offers both *better accuracy* (average closer to the accepted value) and *better precision* (since the value has a lower standard deviation).

1R.24. Using the data provided, the averages and their deviations are as follows:

Data point	Your Data	deviation	Partner's Data	deviation
1	321	–4	327	–0.5
2	326	1	329	1.5
3	324	–1	328	0.5
4	329	4	326	–1.5
Averages:	325		327.5	

Note that the deviations for both methods are calculated by first determining the average of the four data points, and then subtracting the individual data points from the average.

(a) The average mass of aspirin for method A is 325 mg while the average mass of aspirin for method B is 327.5 mg = 328 mg (three sf).

(b) Your data is more accurate than your partner's data as your average mass coincides exactly with the manufacturer's accepted mass. Percent errors shown below.

Error = experimental value – accepted value

From Method A error = (325 – 325) = 0 mg

From Method B error = (327.5 – 325) = 2.5 mg

and the percent error is then:

$$(\text{Method A}) = \frac{0}{325} \times 100\% = 0\%$$

$$(\text{Method B}) = \frac{2.5}{325} \times 100\% = 0.8\% \text{ (1 sf)}$$

(c) Your partner's data is more precise. The data for your partner are closer to each other.

(d) The standard deviation for you and your partner:

Your data:

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$$\sqrt{\frac{(-4)^2 + (1)^2 + (-1)^2 + (4)^2}{3}} = \sqrt{\frac{34}{3}} = 3 \text{ mg (1 sf)}$$

Your partner's data:

$$\sqrt{\frac{(-0.5)^2 + (1.5)^2 + (0.5)^2 + (-1.5)^2}{3}} = \sqrt{\frac{5.0}{3}} = 1 \text{ mg (1 sf)}$$

The smaller standard deviation for your partner's data confirms the prediction.

Exponential Notation and Significant Figures

1R.25. Express the following numbers in exponential (or scientific) notation:

(a) $0.0830 \text{ g} = 8.30 \times 10^{-2} \text{ g}$ To locate the decimal behind the first non-zero digit, we move the decimal place to the right by 2 spaces (–2); 3 significant figures

(b) $136 \text{ g} = 1.36 \times 10^{2} \text{ g}$ To locate the decimal behind the first non-zero digit, we move the decimal place to the left by 2 spaces (+2); 3 significant figures

(c) $0.00602 \text{ g} = 6.02 \times 10^{-3} \text{ g}$ To locate the decimal behind the first non-zero digit, we move the decimal place to the right by 3 spaces (–3); 3 significant figures

(d) $3000 \text{ mL} = 3 \times 10^{3} \text{ mL}$ To locate the decimal behind the first non-zero digit, we move the decimal place to the left by 3 spaces (+3); 1 significant figure

1R.26. Express the following numbers in exponential (or scientific) notation:

(a) $1356 \text{ mL} = 1.356 \times 10^{3} \text{ mL}$ To locate the decimal behind the first non-zero digit, we move the decimal place to the left by 3 spaces (+3); 4 significant figures

(b) $0.03042 \text{ L} = 3.042 \times 10^{-2} \text{ L}$ To locate the decimal behind the first non-zero digit, we move the decimal place to the right by 3 spaces (–3); 4 significant figures

(c) $250.0 = 2.500 \times 10^{2}$ To locate the decimal behind the first non-zero digit, we move the decimal place to the left by 2 spaces (+2); 4 significant figures

(d) $120 = 1.2 \times 10^{2}$ To locate the decimal behind the first non-zero digit, we move the decimal place to the left by 2 spaces (+2); 2 significant figures

1R.27. Express the following without using scientific notation:

(a) $5.43 \times 10^{2} \text{ m} = 543 \text{ m}$ (3 significant figures)

(b) $4.306 \times 10^{-2} \text{ L} = 0.04306 \text{ L}$ (4 significant figures)

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(c) 6.20×10^{-4} L = 0.000620 L (3 significant figures) (zero after decimal point is significant)

(d) 8.42×10^{3} mL = 8420 mL (3 significant figures)

1R.28. Express the following without using scientific notation:

(a) 3.25×10^{-2} m = 0.0325 m (3 significant figures)

(b) 4.02×10^{-3} mL= 0.00402 mL (3 significant figures)

(c) 4.2×10^{3} mL = 4200 mL (2 significant figures)

(d) 9.305×10^{4} g = 93050 g (4 significant figures)

1R.29. Perform operations and report answers to proper number of sf:

(a) $(1.52)\left(6.21 \times 10^{-3}\right) = 9.44 \times 10^{-3}$ (3 sf - each term in the product has 3)

(b) $\left(6.217 \times 10^{3}\right) - \left(5.23 \times 10^{2}\right) = 5.694 \times 10^{3} = 5694$ [Convert 5.23×10^{2} to 0.523×10^{3} and subtract, leaving 5.694×10^{3}. With 3 decimal places to the right of the decimal place in both numbers, we can express the difference with 3 decimal places.

(c) $\left(6.217 \times 10^{3}\right) \div \left(5.23 \times 10^{2}\right) = 11.887$ or 11.9 (3 sf)

Recall that in multiplication and division, the result should have the same number of significant figures as the **term** with the fewest significant figures.

(d) $(0.0546)(16.0000)\left[\frac{7.779}{55.85}\right] = 0.121678$ or 0.122 (3 sf)

the same rule applies, as in part (c) above: The first term has 3 sf, the second term 6 sf—yes the zeroes count, and the third term 4 sf.

Dividing the two terms in the last quotient gives an answer with 4 sf. So with 3, 6, and 4 sf in the terms, the answer should have **no more than** 3 sf

1R.30. Provide result of calculation with proper number of significant figures:

In all of these calculations, the answer has 3 sf

(a) 2.44×10^{8} (c) 0.133

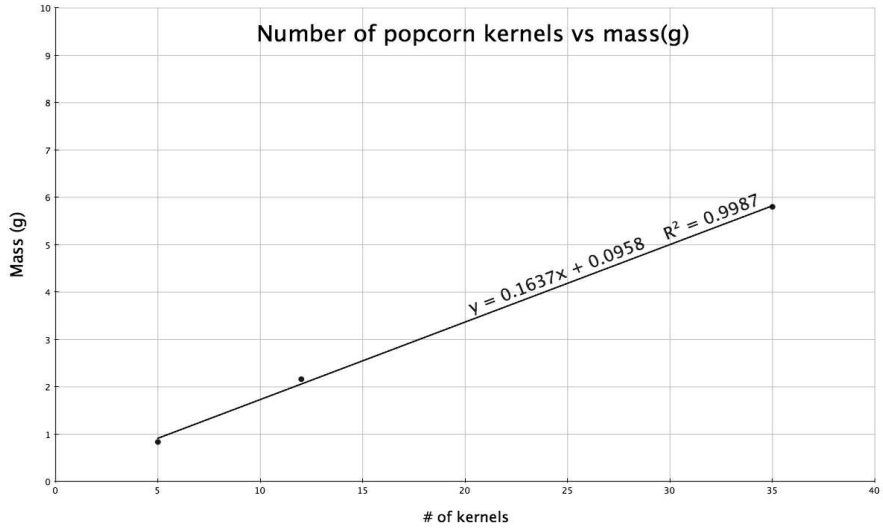
(b) 4.85×10^{-2} (d) 0.0286

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Graphing

1R.31. Plot the data for number of kernels of popcorn versus mass (in grams):



The best straight line has the equation, $y = 0.1637x + 0.096$, with a slope of 0.1637. This slope indicates that the mass increases by a factor of 0.1637grams with each kernel of popcorn. The mass of 20 kernels would be:

mass = (0.1637)(20) + 0.096 or 3.370 grams.

To determine the number of kernels (x) with a mass of 20.88 grams, substitute 20.88 for mass (i.e., y) and solve for the number of kernels.

20.88 g = 0.1637(x) + 0.096; (20.88 – 0.096) = 0.1637x and dividing by the slope:

$$x = \frac{20.88 - 0.096}{0.1637} = 127 \text{ kernals}$$

1R.32. Using the provided graph:

(a) The value of x when $y = 4.0$ is 0.21

(b) The value of y when x is 0.30 is 5.6

(c) slope = $\frac{5.6 - 4.0}{0.30 - 0.21} = 18$; y-intercept = 0.20

(d) The value of y when $x = 1.0$: $y = 18x + 0.20 = (18)(1.0) + 0.20 = 18$

1R.33. Using the graph shown, determine the values of the equation of the line:

(a) Using the points (0.00, 20.00) and (5.00, 0.00), we calculate the slope (rise/run):

$$\frac{20.00 - 0.00}{0.00 - 5.00} = -4.00$$

The y-intercept (b) is the y value when the x value is zero (0).

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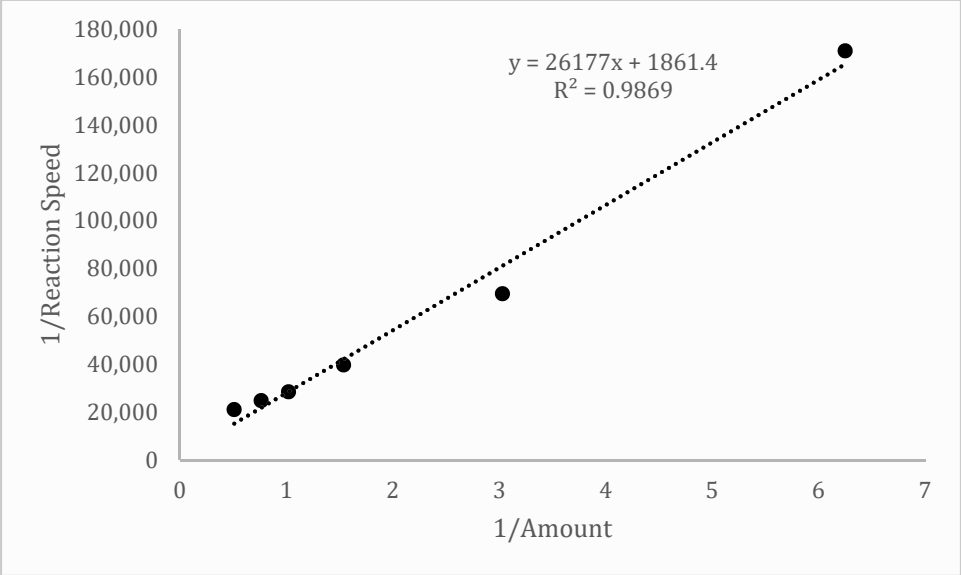
We can read the value of b from the graph (20.00).

The equation for the line is: $y = -4.00x + 20.00$

(b) The value of y when $x = 6.0$ is: $y = (-4.00)(6.0) + 20.00 = -4.00$.

1R.34. (a) The data (and the reciprocals) are as follows:

Amount of H_2O_2	1/Amount	Reaction Speed	1/Speed
1.96	0.5102	4.75×10^{-5}	21050
1.31	0.7634	4.03×10^{-5}	24810
0.98	1.02	3.51×10^{-5}	28490
0.65	1.54	2.52×10^{-5}	39680
0.33	3.03	1.44×10^{-5}	69440
0.16	6.25	0.585×10^{-5}	170900



(b) $y = 2.62 \times 10^4\, x + 1900$ with a slope of 2.62×10^4 (3 sf) and y-intercept = 1900 (rounded to the nearest hundred).

Solving Equations

1R.35. Solving the equation for "C":

$(0.502)(123) = (750.)C$ and rearranging the equation by dividing by 750. gives

$$C = \frac{(0.502)(123)}{750.} = 0.0823 \text{ (3 sf)}$$

1R.36. Solving the equation for n:

$(1.0)(22.4) = n(0.082057)(273.15)$ Rearranging: $n = \dfrac{(1.0)(22.4)}{(0.082057)(273.15)} = 1.0$ (2 sf)

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1R.37. Solve the following equation for T:

$(4.184)(244)(T - 292.0) + (0.449)(88.5)(T - 369.0) = 0$

$020.9T - 298101.6 + 39.74T - 14663 = 0$

$1060.64T - 312764 = 0$; Solving for T:

$$T = \frac{312764}{1060.64} = 295 \text{ (3 sf)}$$

1R.38. Solving the equation for n:

$$-246.0 = 1312\left[\frac{1}{2^2} - \frac{1}{n^2}\right] \text{ Simplify by dividing by 1312: } \frac{-246.0}{1312} = \left[\frac{1}{2^2} - \frac{1}{n^2}\right]$$

$$-0.1875 = \left[\frac{1}{4} - \frac{1}{n^2}\right] \text{ Substituting 0.25 for 1/4, } -0.1875 = 0.25 - \frac{1}{n^2}$$

$\frac{1}{n^2} = 0.4375$; Multiply by n^2, divide by 0.4375 and take the square root of both sides to obtain: $n = \sqrt{\frac{1}{0.4375}} = 2$ (1 sf)

General Questions

1R.39. Express the length 1.97 Angstroms in nanometers; In picometers;

$$\frac{1.97 \text{ Å}}{1} \cdot \frac{1 \times 10^{-10} \text{ m}}{1 \text{ Å}} \cdot \frac{1 \times 10^{9} \text{ nm}}{1 \text{ m}} = 0.197 \text{ nm}$$

$$\frac{1.97 \text{ Å}}{1} \cdot \frac{1 \times 10^{-10} \text{ m}}{1 \text{ Å}} \cdot \frac{1 \times 10^{12} \text{ pm}}{1 \text{ m}} = 197 \text{ pm}$$

1R.40. Separation between C atoms in diamond is 0.154 nm.

(a) In meters: $\frac{0.154 \text{ nm}}{1} \cdot \frac{1 \text{ m}}{1 \times 10^{9} \text{ nm}} = 1.54 \times 10^{-10} \text{ m}$

(b) In picometers: $1.54 \times 10^{-10} \text{ m} \cdot \frac{1 \text{ pm}}{1 \times 10^{-12} \text{ m}} = 154 \text{ pm}$

(c) In Angstroms: $1.54 \times 10^{-10} \text{ m} \cdot \frac{1 \text{ Å}}{1 \times 10^{-10} \text{ m}} = 1.54 \text{ Å}$

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1R.41. Diameter of red blood cell = 7.5 μm

(a) In meters: $\frac{7.5\ \mu\text{m}}{1} \cdot \frac{1\ \text{m}}{1 \times 10^{6}\ \mu\text{m}} = 7.5 \times 10^{-6}\ \text{m}$

(b) In nanometers: $\frac{7.5\ \mu\text{m}}{1} \cdot \frac{1\ \text{m}}{1 \times 10^{6}\ \mu\text{m}} \cdot \frac{1 \times 10^{9}\ \text{nm}}{1\ \text{m}} = 7.5 \times 10^{3}\ \text{nm}$

(c) In picometers: $\frac{7.5\ \mu\text{m}}{1} \cdot \frac{1\ \text{m}}{1 \times 10^{6}\ \mu\text{m}} \cdot \frac{1 \times 10^{12}\ \text{m}}{1\ \text{m}} = 7.5 \times 10^{6}\ \text{pm}$

1R.42. $1.53\ \text{g cisplatin} \cdot \frac{65.0\ \text{g Pt}}{100.0\ \text{g cisplatin}} = 0.995\ \text{g Pt}$

1R.43. Mass of procaine hydrochloride (in mg) in 0.50 mL of solution

$$\frac{0.50\ \text{mL}}{1} \cdot \frac{1.0\ \text{g}}{1\ \text{mL}} \cdot \frac{10.\ \text{g procaine HCl}}{100\ \text{g solution}} \cdot \frac{10^{3}\ \text{mg procaine HCl}}{1\ \text{g procaine HCl}} = 50.\ \text{mg procaine HCl}$$

1R.44. Length of a cube of Al to have a mass of 7.6 grams:

$7.6\ \text{g} \cdot \frac{1\ \text{cm}^3}{2.698\ \text{g}} = 2.8\ \text{cm}^3$; The cube has dimensions of the length^3.

The cube edge = $\sqrt[3]{2.8\ \text{cm}^3} = 1.4\ \text{cm}$

1R.45. Average density of a marble: $\frac{95.2\ \text{g}}{(99 - 61)\ \text{mL}} \cdot \frac{1\ \text{mL}}{1\ \text{cm}^3} = 2.5\ \text{g/cm}^3$

1R.46. Calculate the density of the solid and identify the compound:

$$\frac{18.82\ \text{g}}{(15.3 - 8.5)\ \text{mL}} \cdot \frac{1\ \text{mL}}{1\ \text{cm}^3} = 2.8\ \text{g/cm}^3$$

The white solid's density matches that of (c) KBr.

1R.47. For the sodium chloride unit cell:

(a)The volume of the unit cell is the (edge length)^3

With an edge length of 0.563 nm, the volume is (0.563 nm)^3 or 0.178 nm^3. The volume in cm^3 is calculated by first expressing the edge length in cm: ,

$\frac{0.563\ \text{nm}}{1} \cdot \frac{1 \times 10^{2}\ \text{cm}}{1 \times 10^{9}\ \text{nm}} = 5.63 \times 10^{-8}\ \text{cm}$, so the volume is $1.78 \times 10^{-22}\ \text{cm}^3$

(b) The mass of the unit cell is:

$M = V \cdot D = 1.78 \times 10^{-22}\ \text{cm}^3 \cdot 2.17\ \text{g/cm}^3 = 3.87 \times 10^{-22}\ \text{g}$

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(c) Given the unit cell contains 4 NaCl formula units, the mass of one formula unit is:

$$\frac{3.87 \times 10^{-22} \text{ g per unit cell}}{4 \text{ formula units per unit cell}} = 9.68 \times 10^{-23} \text{ g/formula unit}$$

1R.48. Volume of a 1.50-carat diamond:

$$1.50 \text{ carat} \cdot \frac{0.200 \text{ g}}{1 \text{ carat}} \cdot \frac{1 \text{ cm}^3}{3.513 \text{ g}} = 0.0854 \text{ cm}^3$$

1R.49. The accepted value for a normal human temperature is 98.6 ˚F. On the Celsius scale this corresponds to: $^{\circ}\text{C} = \frac{5 \text{ °C}}{9 \text{ °F}}(98.6 \text{ °F} - 32 \text{ °F}) = 37 \text{ °C}$

Since the melting point of gallium is 29.8 ˚C, the gallium should melt in your hand.

1R.50. A good estimate for the density would be between that at 15 °C and that at 25 °C. If one assumes it is halfway between those densities, a better estimate would be 0.99810 g/cm^3. It should be noted, however, that the decrease in density with increasing temperature is not linear.

1R.51. The heating of popcorn causes the loss of water.

(a) The percentage of mass lost upon popping:

$$\frac{(0.125 \text{ g} - 0.106 \text{ g})}{0.125 \text{ g}} \times 100\% = 15\%$$

(b) With an average mass of 0.125 g, the number of kernels in a pound of popcorn:

$$\frac{1 \text{ kernel}}{0.125 \text{ g}}\left(\frac{453.6 \text{ g}}{1 \text{ lb}}\right) = 3630 \text{ kernels}$$

1R.52. What is the thickness of the aluminum foil in millimeters?

$$\text{Volume} = \frac{12 \text{ oz}}{1} \cdot \frac{28.4 \text{ g}}{1 \text{ oz}} \cdot \frac{1 \text{ cm}^3}{2.70 \text{ g}} = 126 \text{ cm}^3$$

$$\text{Area} = 75 \text{ ft}^2\left(\frac{12 \text{ in}}{1 \text{ ft}}\right)^2\left(\frac{2.54 \text{ cm}}{1 \text{ in}}\right)^2 = 7.0 \times 10^4 \text{ cm}^2$$

$$\text{Thickness} = \frac{\text{volume}}{\text{area}} = \frac{126 \text{ cm}^3}{7.0 \times 10^4 \text{ cm}^2} = 1.8 \times 10^{-3} \text{ cm} = 1.8 \times 10^{-2} \text{ mm}$$

1R.53. The mass of NaF needed for 150,000 people for a year:

One way to do this problem is to begin with a factor that contains the units of the "answer" (mass NaF in kg). Since NaF is 45% fluoride (and 55 %Na), we can write the factor:

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$\frac{100.0\text{ kg NaF}}{45.0\text{ kg F}^-}$. Note that the expression of kg/kg has the same value as the expression g/g—and provides the "desired" units of our answer.

A concentration of 1 ppm can be expressed as: $\frac{1.00\text{ kg F}^-}{1.00 \times 10^6\text{ kg H}_2\text{O}}$ [We could use the fraction with the masses expressed in **grams**, but we would have to convert grams to kg. Note this factor can be derived from the factor using grams if you multiply BOTH numerator and denominator by 1000.]. Using the data provided in the problem, plus conversion factors (found in the rear cover of your textbook)

$$150{,}000\text{ people} \cdot \frac{660\text{ L H}_2\text{O}}{1\text{ person} \cdot \text{day}} \cdot \frac{365\text{ days}}{\text{year}} \cdot \frac{1.00\text{ kg H}_2\text{O}}{1.00\text{ L H}_2\text{O}} \cdot \frac{1.00\text{ F}^-}{1.00 \times 10^6\text{ kg H}_2\text{O}} \cdot \frac{100\text{ kg NaF}}{45.0\text{ kg F}^-} = 8.0 \times 10^4\text{ kg NaF/day}$$

Note that 660 L of water per day limits the answer to 2 sf.

1R.54. Area in cm^2: $0.5\text{ acre} \cdot \frac{1.0 \times 10^4\text{ m}^2}{2.47\text{ acres}}\left(\frac{100\text{ cm}}{1\text{ m}}\right)^2 = 2 \times 10^7\text{ cm}^2$

$$\text{thickness} = \frac{\text{volume}}{\text{area}} = \frac{5\text{ cm}^3}{2 \times 10^7\text{ cm}^2} = 2 \times 10^{-7}\text{ cm}$$

This is likely related to the "length" of oil molecules.

1R.55. Mass of sulfuric acid in 750. mL (or 750. cm^3) solution.

$$750.\text{ cm}^3\text{ solution} \cdot \frac{1.285\text{ g solution}}{1\text{ cm}^3\text{ solution}} \cdot \frac{38.08\text{ g sulfuric acid}}{100\text{ g solution}} = 367\text{ g sulfuric acid (3 sf)}$$

1R.56. Mass of hydrochloric acid in 50. mL of 37% by mass solution.

$$50.\text{ mL HCl solution}\left(\frac{1.2\text{ g HCl solution}}{1\text{ mL solution}}\right) = 60.\text{ g HCl solution}$$

$$60.\text{ g HCl solution}\left(\frac{37\text{ g HCl}}{100\text{ g HCl solution}}\right) = 22\text{ g HCl (2sf)}$$

1R.57. Volume of 5.25% bleach to contain 8.0 g of NaClO.

$$8.0\text{ g NaClO} \cdot \frac{100\text{ g bleach}}{5.25\text{ g NaClO}} \cdot \frac{1\text{ mL bleach}}{1.1\text{ g bleach}} = 140\text{ mL bleach (2 sf)}$$

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1R.58. The volume is: $\frac{279\ \text{kg}}{1} \cdot \frac{10^3\ \text{g}}{1\ \text{kg}} \cdot \frac{1\ \text{cm}^3}{19.3\ \text{g}} = 1.45 \times 10^4\ \text{cm}^3$

volume = (area)(thickness)

$$1.45 \times 10^4\ \text{cm}^3 = (\text{area}) \left(0.0015\ \text{mm} \times \frac{1\ \text{cm}}{10\ \text{mm}} \right)$$

Solve for area:

$$\text{area} = \frac{1.45 \times 10^4\ \text{cm}^3}{\left(0.0015\ \text{mm} \cdot \frac{1\ \text{cm}}{10\ \text{mm}} \right)} = 9.6 \times 10^7\ \text{cm}^2$$

Converting to m^2 : $9.6 \times 10^7\ \text{cm}^2 \left(\frac{1\ \text{m}}{100\ \text{cm}} \right)^2 = 9.6 \times 10^3\ \text{m}^2$

1R.59. Density of water changes with state:

(a) Volume of solid water at –10°C when a 250. mL can is filled with liquid water at 25°C:

The volume of liquid water at 25 °C is 250. mL (or cm^3). The mass of that water is:

$$\frac{0.997\ \text{g water}}{1.000\ \text{cm}^3\ \text{water}} \cdot \frac{250.\ \text{cm}^3}{1} = 249.25\ \text{g water (249 to 3 sf)}$$

That mass of water at the lower temperature will occupy:

$$\frac{1.000\ \text{cm}^3\ \text{water}}{0.917\ \text{g water}} \cdot \frac{249.25\ \text{g water}}{1} = 271.81\ (\text{or } 272\ \text{cm}^3\ \text{to 3 sf})$$

(b) With the can being filled to 250. mL at room temperature, the expansion (an additional 22 mL) cannot be contained in the can. (Get out the sponge—there's a mess to clean up.)

1R.60. Volume of room and mass of air:

(a) Volume of room = (length)(width)(height)

$$= (18\ \text{ft})(15\ \text{ft})(8.5\ \text{ft}) \left(\frac{12\ \text{in}}{1\ \text{ft}} \right)^3 \left(\frac{2.54\ \text{cm}}{1\ \text{in}} \right)^3 \left(\frac{1\ \text{m}}{100\ \text{cm}} \right)^3 = 65\ \text{m}^3$$

$$\text{and } 65\ \text{m}^3 \cdot \frac{1\ \text{L}}{10^{-3}\ \text{m}^3} = 6.5 \times 10^4\ \text{L}$$

(b) Mass of air in kg: $\frac{6.5 \times 10^4\ \text{L}}{1} \cdot \frac{1.2\ \text{g}}{1\ \text{L}} \cdot \frac{1\ \text{kg}}{10^3\ \text{g}} = 78\ \text{kg}$

and in pounds: $\frac{78\ \text{kg}}{1} \cdot \frac{10^3\ \text{g}}{1\ \text{kg}} \cdot \frac{1\ \text{lb}}{454\ \text{g}} = 170\ \text{lb}$

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1R.61. Calculate the density of steel if a steel sphere of diameter 9.40 mm has a mass of 3.475 g:

The radius of the sphere is (0.5)(9.40 mm) or 4.70 mm. Since density is usually expressed in cm^3, express the radius in cm (0.470 cm) and substitute into the volume equation:

$$V=\frac{4}{3}\pi r^3 = \frac{4}{3}(3.1416)\left(0.470\text{cm}\right)^3 = 0.435\ \text{cm}^3 \text{(to 3 sf)}$$

The density is 3.475 g /0.435 cm^3 = 7.99 g/cm^3.

1R.62. Identify the liquid:

(a) density = $\dfrac{16.08\ \text{g} - 12.20\ \text{g}}{3.50\ \text{mL}} \cdot \dfrac{1\ \text{mL}}{1\ \text{cm}^3} = 1.11\ \text{g/cm}^3$.

Of the liquids supplied, the liquid is most likely ethylene glycol.

(b) With 2 sf, the calculated density would be 1.1 g/cm^3. While this value still suggests that the liquid is ethylene glycol, it is close to the values for acetic acid and water. Additional testing would be needed to uniquely identify the liquid.

1R.63. (a) Calculate the density of an irregularly shaped piece of metal:

$$\text{Density} = \frac{m}{V} = \frac{74.122\ \text{g}}{36.7\ \text{cm}^3 - 28.2\ \text{cm}^3} = \frac{74.122\ \text{g}}{8.5\ \text{cm}^3} = 8.7\ \text{g/cm}^3$$

Note that the subtraction of volumes leaves only 2 sf, limiting the density to 2 sf

(b) From the list of metals provided, one would surmise that the metal is **cadmium**. Since the major uncertainty is in the volume, one can substitute 8.4 and 8.6 cm^3 as the volume and calculate the density (resulting in 8.82 and 8.62 g/cm^3 respectively). The hypothesis that the metal is cadmium is reasonably sound, but the calculated density is close to that of cobalt, nickel, and copper. Further testing should be done on the metal.

1R.64. (a) Density = $\dfrac{3.2745\ \text{g}}{5.0\ \text{mL}} = 0.65\ \text{g/mL}$

(b) Using volumes of 4.9 mL and 5.1 mL, corresponding to ±0.1 mL of the recorded volume, the range of densities possible is 0.66 g/mL to 0.64 g/mL. No, all five hydrocarbon density values fall within the range of possible values for the liquid.

(c) No, all five hydrocarbon density values fall within the range of possible values.

(d) Using a volume of (4.93 ± 0.01) mL, the calculated density is 0.664 g/mL, and maximum and minimum density values are 0.666 g/mL and 0.663 g/mL. The liquid is 2-methylpentane.

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1R.65. Mass of Hg in the capillary:

Mass of capillary with Hg	3.416 g
Mass of capillary without Hg	3.263 g
Mass of Hg	0.153 g

To determine the volume of the capillary, calculate the volume of Hg that is filling it.

$$0.153 \text{ g Hg} \cdot \frac{1 \text{ cm}^3}{13.546 \text{ g Hg}} = 1.13 \times 10^{-2} \text{ cm}^3 \text{ (3 sf)}$$

Now that we know the volume of the capillary, and the length of the tubing (given as 16.75 mm—or 1.675 cm), we can calculate the radius of the capillary using the equation: Volume = $\pi r^2 l$.

1.13×10^{-2} cm^3 = (3.1416) r^2(1.675 cm), and solving for r^2:

$$r^2 = \frac{1.13 \times 10^{-2} \text{ cm}^3}{(3.1416)(1.675 \text{ cm})} = 2.15 \times 10^{-3} \text{ cm}^2 \text{ and } r = 4.63 \times 10^{-2} \text{ cm}.$$

Diameter = $2r = 2(4.63 \times 10^{-2} \text{ cm}) = 0.0927$ cm

1R.66. Volume of this amount of copper: $57 \text{ kg} \cdot \frac{10^3 \text{ g}}{1 \text{ kg}} \cdot \frac{1 \text{ cm}^3}{8.96 \text{ g}} = 6.4 \times 10^3 \text{ cm}^3$ copper

$$\text{The length of wire} = \frac{\text{volume}}{(\pi)(\text{radius})^2} = \frac{6.4 \times 10^3 \text{ cm}^3}{(\pi)(0.950 \text{ cm}/2)^2} = 9.0 \times 10^3 \text{ cm}$$

and expressed in meters: $9.0 \times 10^3 \text{ cm} \cdot \frac{1 \text{ m}}{100 \text{ cm}} = 90. \text{ M}$

1R.67. Regarding a cube of Cu:

(a) The number of Cu atoms in a cube whose mass is 0.1206 g.

$$\frac{0.1206 \text{ g}}{\text{cube}} \cdot \frac{1 \text{ atom Cu}}{1.055 \times 10^{-22} \text{g}} = 1.143 \times 10^{21} \text{ atoms Cu}$$

Fraction of the lattice that contains Cu atoms:

Given the radius of a Cu atom is 128 pm, and the number of Cu atoms, the total volume occupied by the Cu atoms is the volume occupied by ONE atom ($4/3\pi r^3$) multiplied by the total number of atoms:

Volume of one atom: $4/3 \cdot 3.1416 \cdot (128\text{pm})^3 = 8.78 \times 10^6 \text{ pm}^3$

Total volume:$(8.78 \times 10^6 \text{ pm}^3/\text{Cu atom})(1.143 \times 10^{21} \text{ atoms Cu})$

$= 1.00 \times 10^{28} \text{ pm}^3$.

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The lattice cube has a volume of $(0.236\ \text{cm})^3$ or $(2.36 \times 10^9\ \text{pm})^3$ or $1.31 \times 10^{28}\ \text{pm}^3$.

The fraction occupied is: total volume of Cu atoms/total volume of lattice cube:

$$\frac{1.00 \times 10^{28}\ \text{pm}^3}{1.31 \times 10^{28}\ \text{pm}^3} = 0.764 \text{ or } 76.4\% \text{ occupied (3 sf)}$$

The empty space in a lattice is due to the inability of spherical atoms to totally fill a given volume. A macroscopic example of this phenomenon is visible if you place four marbles in a square arrangement. At the center of the square there are voids. In a cube, there are similar repeating incidents.

(b) Estimate the number of Cu atoms in the smallest repeating unit:

Since we know the length of the smallest repeating unit (the unit cell), let's calculate the volume (first converting the length to units of centimeters:

$$\text{L} = 361.47 \text{ pm so } \text{L} = 361.47\ \text{pm} \cdot \frac{1 \times 10^2\ \text{cm}}{1 \times 10^{12}\ \text{pm}} = 3.6147 \times 10^{-8}\ \text{cm}$$

$$\text{V} = \text{L}^3 = (3.6147 \times 10^{-8}\ \text{cm})^3 = 4.7230 \times 10^{-23}\ \text{cm}^3$$

Since we know the density, we can calculate the mass of one unit cell:

$\text{D x V} = 8.960\ \text{g/ cm}^3 \times 4.7230 \times 10^{-23}\ \text{cm}^3 = 4.232 \times 10^{-22}\ \text{g}$

Knowing the mass of one copper atom (1.055×10^{-22} g) we can calculate the number of Cu atoms in that mass:

$$\frac{4.232 \times 10^{-22}\ \text{g}}{1.055 \times 10^{-22}\ \text{g/Cu atom}} = 4.011 \text{ Cu atoms}$$

Four atoms are present in this smallest repeating unit. As you will learn later, the number of atoms for a face-centered cubic lattice is 4.

1R.68. Determine the density of lead, average density, percent error, and standard deviation using the provided data:

Data	Density (g/cm^3)
1	11.6
2	11.8
3	11.5
4	12.0
Average	11.7

$$\text{Percent error} = \frac{\text{error in measurement}}{\text{accepted value}} \times 100\% = \frac{11.7\ \text{g/cm}^3 - 11.3\ \text{g/cm}^3}{11.3\ \text{g/cm}^3} \times 100\% = 4\%\ (1\ \text{sf})$$

Standard Deviation is calculated as usual:

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Data	Density (g/cm^3)	Difference (Measurement – Average)	Square of Difference
1	11.6	– 0.1	1×10^{-2}
2	11.8	0.1	1×10^{-2}
3	11.5	– 0.2	4×10^{-2}
4	12.0	0.3	9×10^{-2}

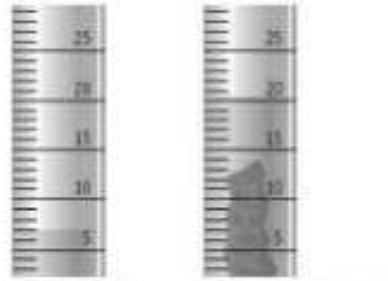
The sum of squares is then 0.15 and square root of $0.15/3 = 0.2$ g/cm^3

In The Laboratory

1R.69. The metal will displace a volume of water that is equal to the volume of the metal.

The difference in volumes of water (20.2 – 6.9) corresponds to the volume of metal. Since 1 mL = 1 cm^3, the density of the metal is then:

$$\frac{\text{Mass}}{\text{Volume}} = \frac{37.5\text{ g}}{13.3\text{ cm}^3} = \text{or } 2.82\ \frac{\text{g}}{\text{cm}^3}$$



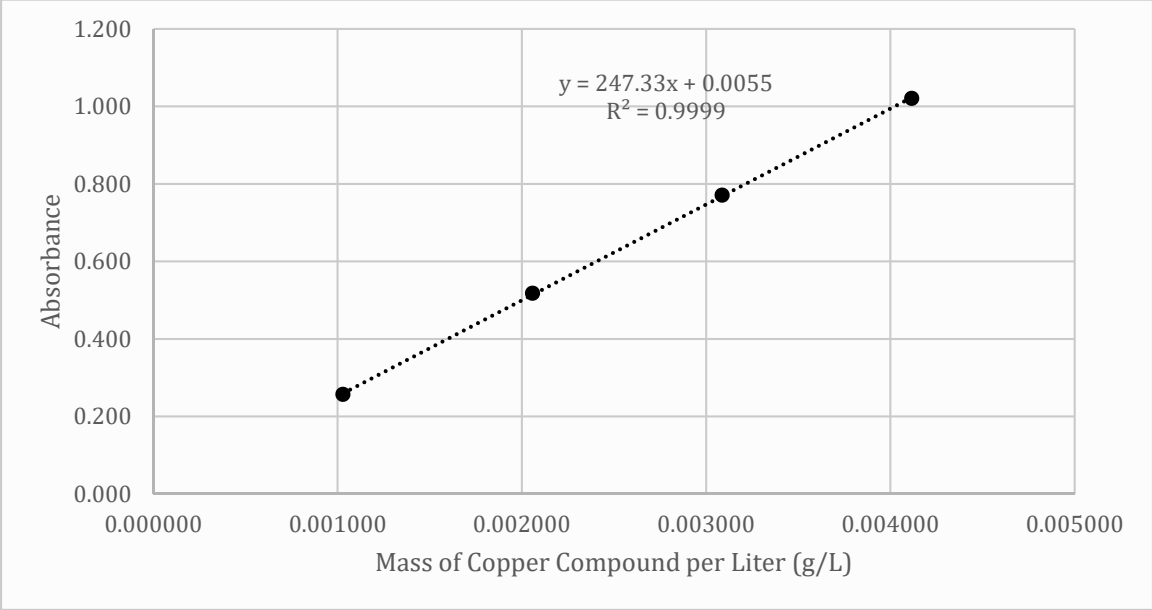
Graduated cylinders with unknown metal (right)

The metal is (d) aluminum.

1R.70. Calculate the density of the sample: $\frac{23.5\text{ g}}{(52.2 - 47.5)\text{ mL}} \cdot \frac{1\text{ mL}}{1\text{ cm}^3} = 5.0\ \text{g/cm}^3$

The sample's density matches that of fool's gold.

1R.71. Plot of Absorbance vs mass of compound (g/L):



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The equation for the best straight line is $y = 247x + 0.006$

The best straight line indicates that the slope is 247 and the intercept is 0.006

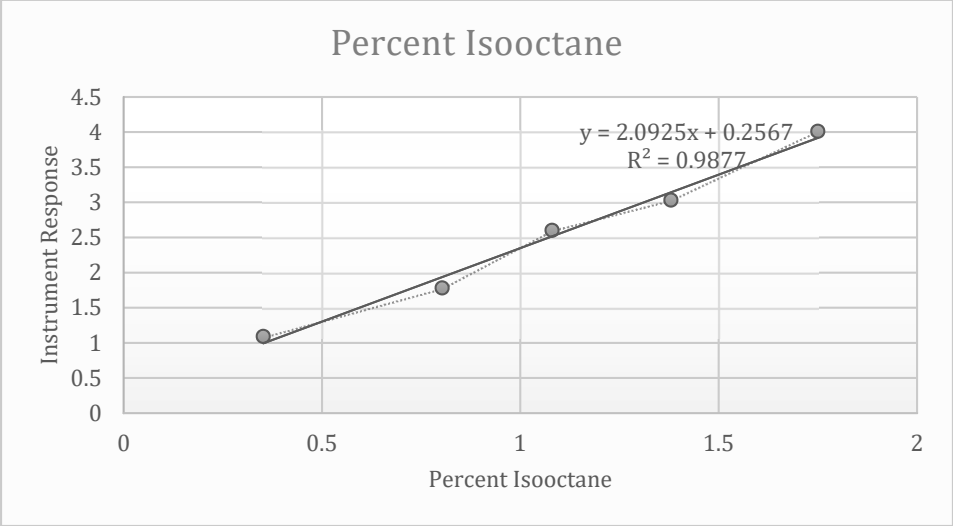
If Absorbance is 0.635, what is the mass of Cu in g/L and mg/mL:

Substituting into the equation for the line:

$0.635 = (247)(\text{concentration}) + 0.006$; concentration $= 2.55 \times 10^{-3}$ g/L

$$\frac{2.55 \times 10^{-3}\ \text{g}}{\text{L}} \cdot \frac{1\ \text{L}}{10^{3}\ \text{mL}} \cdot \frac{10^{3}\ \text{mg}}{1\ \text{g}} = 2.55 \times 10^{-3}\ \text{mg/mL}$$

1R.72. Use data to determine % isooctane:



The best straight line indicates that the slope is 2.09 and the intercept is 0.26.

Substituting into the equation for the line with instrument response = 2.75:

2.75 = 2.09(% isooctane) + 0.26 and solving for (% isooctane)

% isooctane = 1.19%

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1R.73. Insert the data in a spreadsheet (here, Excel is used), to obtain the results:

Student	% Acetic Acid
1	5.22
2	5.28
3	5.22
4	5.30
5	5.19
6	5.23
7	5.33
8	5.26
9	5.15
10	5.22
Average	5.24
St.Dev	0.05

Only the values 5.30%, 5.33%, and 5.15% fall *outside* the Average ± St.Dev, so seven points fall within the range $5.19 \leq x \leq 5.29$.

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Solution and Answer Guide

Kotz Treichel Townsend Treichel, Chemistry and Chemical Reactivity 11e, 978-0-357-85140-1, Chapter 2: Atoms, Molecules, and Ions

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Applying Chemical Principles:

Using Isotopes

2.1.1. Neutrons in ^{18}O; Neutrons in the two Pb isotopes:

Neutrons in ^{18}O: atomic mass (18) – proton number (8) = 18 – 8 = 10 neutrons

Neutrons in ^{206}Pb: atomic mass (206) – proton number (82) = 206 – 82 = 124 neutrons

Neutrons in ^{204}Pb: atomic mass (204) – proton number (82) = 204 – 82 = 122 neutrons

2.1.2. Using the abundance and mass of the three stable isotopes of O, calculate the atomic weight:

Atomic weight is a weighted average of the various isotopes. Add the product of (percent abundance) (isotopic mass) for all the isotopes to obtain the atomic weight:

Isotopic Mass	% Abundance (converted to a fraction)	Product
15.9949	0.99757	15.956
16.9991	0.00038	0.0065
17.9992	0.00205	0.0369
	SUM:	15.999

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Arsenic, Medicine, and the Formula of Compound 606

2.2.1. What is the empirical formula of the compound enargite?

The percentages total to 100%, so think of the percent of each element in 100 g of the compound: Mass As: 19.024 g; Mass Cu: 48.407 g; Mass S: 32.569 g.
Now convert the masses to moles of each element:

$$19.024 \text{ g As}\left(\frac{1 \text{ mol As}}{74.922 \text{ g As}}\right) = 0.25392 \text{ mol As}$$

$$48.407 \text{ g Cu}\left(\frac{1 \text{ mol Cu}}{63.546 \text{ g Cu}}\right) = 0.76176 \text{ mol Cu}$$

$$32.569 \text{ g S}\left(\frac{1 \text{ mol S}}{32.06 \text{ g S}}\right) = 1.016 \text{ mol S}$$

Calculate the mole: mole ratio:

[Hint: divide the amounts (moles) by the smallest amount.]

$$\frac{1.016 \text{ mol S}}{0.25392 \text{ mol As}} = 4.001 \quad \frac{0.76176 \text{ mol Cu}}{0.25392 \text{ mol As}} = 3.0000$$

The ratio of the three elements is: $Cu_3As_1S_4$ or usually written: Cu_3AsS_4.

2.2.2. Molecular formulas of the compounds composing Salvarsan:

First determine empirical formulas for the substances:

$$39.37 \text{ g C}\left(\frac{1 \text{ mol C}}{12.011 \text{ g C}}\right) = 3.278 \text{ mol C}$$

$$3.304 \text{ g H}\left(\frac{1 \text{ mol H}}{1.008 \text{ g H}}\right) = 3.278 \text{ mol H}$$

$$8.741 \text{ g O}\left(\frac{1 \text{ mol O}}{15.999 \text{ g O}}\right) = 0.5463 \text{ mol O}$$

$$7.652 \text{ g N}\left(\frac{1 \text{ mol N}}{14.007 \text{ g N}}\right) = 0.5463 \text{ mol N}$$

$$40.932 \text{ g As}\left(\frac{1 \text{ mol As}}{74.922 \text{ g As}}\right) = 0.54633 \text{ mol As}$$

Determine mole:mole ratios:

Using the smallest # of mol (either for O, N, or As), determine ratios of C and H:

$$\frac{3.278 \text{ mol C}}{0.5463 \text{ mol N}} = \frac{6.000 \text{ mol C}}{1 \text{ mol N}} \text{ or 6 mol C/mol N}$$

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Similarly, the ratio for H/N is also 6 mol H/mol N.

These ratios imply an empirical formula that is C_6H_6AsNO.

Determine the empirical formula mass.

$$6 \text{ mol C}\left(\frac{12.011 \text{ g C}}{1 \text{ mol C}}\right)+6 \text{ mol H}\left(\frac{1.008 \text{ g H}}{1 \text{ mol H}}\right)+1 \text{ mol N}\left(\frac{14.007 \text{ g N}}{1 \text{ mol N}}\right)$$

$$+ 1 \text{ mol O}\left(\frac{15.999 \text{ g O}}{1 \text{ mol O}}\right)+1 \text{ mol As}\left(\frac{74.922 \text{ g As}}{1 \text{ mol As}}\right)=183.04 \text{ g/empirical formula}$$

We are told that one of the compounds has a molar mass of 549 g/mole.

$$\frac{549 \text{ g/mol compound}}{183.04 \text{ g/mol empirical formula}} = 3.00 \text{ mol empirical formula/mol compound}$$

So, this compound has a molecular formula that is 3 empirical formulas/molecular compound,

$C_{18}H_{18}As_3N_3O_3$.

The other compound has a molar mass of 915 g/mol. Using a similar procedure to that above, there are 5 empirical formulas/molecular compound, or a molecular formula of $C_{30}H_{30}As_5N_5O_5$.

Argon-An Amazing Discovery

2.3.1. Volume of globe:

$$0.20389 \text{ g N}\left(\frac{1 \text{ L}}{1.25718 \text{ g N}}\right) = 0.16218 \text{ L or } 162.18 \text{ mL or } 162.18 \text{ cm}^3$$

2.3.2. Determine the density of argon:
Using the prescribed procedure of summing the products of
(fractional volume)(density):
$(0.2096)(1.42952 \text{ g/L}) + (0.7811)(1.25092 \text{ g/L}) + (0.00930)x = 1.000(1.29327 \text{ g/L})$
and $x = 1.78$ g/L

2.3.3. From the periodic table, the atomic weight of Argon = 39.95. Now we deduce the abundance of ^{40}Ar. 100% – 0.334% – 0.063% = 99.603%

Now that we know the abundance of ^{40}Ar, we can determine the weighted average of the three isotopes: $(0.00334)(35.967545) + (0.00063)(37.962732) + (0.99603)x = 39.95$
and $x = 39.96$

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2.3.4. Number of Ar atoms in a room 4.0 m × 5.0 m × 2.4 m:

Volume of room: $4.0\text{ m} \cdot 5.0\text{ m} \cdot 2.4\text{ m} \cdot \dfrac{1\text{ L}}{1.00 \times 10^{-3}\text{ m}^3} = 4.8 \times 10^4\text{ L}$

Use this volume and the density of argon:

$$4.8 \times 10^4\text{ L}\left(\frac{1.78\text{ g Ar}}{1\text{ L}}\right)\left(\frac{1\text{ mol Ar}}{39.95\text{ g Ar}}\right)\left(\frac{6.022 \times 10^{23}\text{ atoms Ar}}{1\text{ mol Ar}}\right) = 1.3 \times 10^{27}\text{ atoms Ar}$$

Practicing Skills

Atoms: Their Composition and Structure

2.1.

Fundamental Particles	Protons	Electrons	Neutrons
Electrical Charges	+1	–1	0
Present in nucleus	Yes	No	Yes
Relative Masses	1.007	**0.00055**	1.009

Note that an electron is 1830 times less massive than a proton or neutron.

2.2. Mass number is the sum of the number of protons and number of neutrons for an atom. Atomic mass is the mass of an atom. Mass number is a simple TOTAL, while atomic mass conveys the sum of the weight of the individual particles.

2.3. Begin by expressing the diameter of the nucleus and the electron cloud in the same units.
2mm (diameter of nucleus) = 2×10^{-3} m (since 1 m = 10^3 mm, 1 mm = 10^{-3} m)

The ratio of diameters: $\dfrac{\text{electron cloud}}{\text{nucleus}} = \dfrac{200\text{ m}}{2 \times 10^{-3}\text{ m}}$. So we set the actual diameters in the same ratio: $\dfrac{200\text{ m}}{2 \times 10^{-3}\text{ m}} = \dfrac{1 \times 10^{-8}\text{ cm}}{x}$ and solving for x:

(200 m) x = (2×10^{-3} m)(1×10^{-8} cm) and $x = 1 \times 10^{-13}$ cm.

Note that we left the actual diameter of the electron cloud in units of centimeters, so the **ratio** would be the same as if we had changed the units to **meters**.

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2.4. Each gold atom has a diameter of 2×145 pm = 290. pm

$$36 \text{ cm}\left(\frac{1 \text{ m}}{100 \text{ cm}}\right)\left(\frac{1 \times 10^{12} \text{ pm}}{1 \text{ m}}\right) = 3.6 \times 10^{11} \text{ pm}$$

$$3.6 \times 10^{11} \text{ pm}\left(\frac{1 \text{ Au atom}}{290. \text{ pm}}\right) = 1.2 \times 10^{9} \text{ Au atoms}$$

2.5. Isotopic symbol for:

(a) K (at. no. 19) with 22 neutrons: 19 + 22 = 41 $^{41}_{19}\text{K}$

(b) V (at. no. 23) with 28 neutrons: 23 + 28 = 51 $^{51}_{23}\text{V}$

(c) Ga (at. no. 31) with 38 neutrons: 31 + 38 = 69 $^{69}_{31}\text{Ga}$

The mass number represents the SUM of the protons + neutrons in the nucleus of an atom. The atomic number represents the # of protons, so (atomic no. + # neutrons) = mass number

2.6. Isotopic symbol for:

(a) Fe (at. no. 26) with 30 neutrons: 26 + 30 = 56 $^{56}_{26}\text{Fe}$

(b) U (at. no. 92) with 146 neutrons: 92 + 146 = 238 $^{238}_{92}\text{U}$

(c) Pb (at. no. 82) with 125 neutrons: 82 + 125 = 207 $^{207}_{82}\text{Pb}$

2.7.

substance	protons	neutrons	electrons
(a) magnesium-24	12	12	12
(b) tin-119	50	69	50
(c) thorium-232	90	142	90
(d) carbon-13	6	7	6
(e) copper-63	29	34	29
(f) bismuth-205	83	122	83

Note that the number of protons and electrons are **equal** for any **neutral atom**. The number of protons is **always** equal to the atomic number. The mass number equals the sum of the numbers of protons and neutrons.

2.8. (a) Number of protons = number of electrons = 43; number of neutrons = 56
(b) Number of protons = number of electrons = 95; number of neutrons = 146

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2.9. Cobalt has an atomic number of 27 that equals the proton and electron number. Remember the mass number is the sum of the protons and neutrons. The three mass numbers are:

27 + 30 = 57, 27 + 31 = 58, 27 + 33 = 60

Thus, we have the following three complete symbols $^{57}_{27}\text{Co}$, $^{58}_{27}\text{Co}$, $^{60}_{27}\text{Co}$

2.10. Atomic number of Ag is 47; both isotopes have 47 protons and 47 electrons.

^{107}Ag 107 – 47 = 60 neutrons and ^{109}Ag 109 – 47 = 62 neutrons

2.11. Hydrogen has three isotopes:

Name	# Protons	# Neutrons	# Electrons
Protium	1	0	1
Deuterium	1	1	1
Tritium	1	2	1

The ONLY difference between the isotopes of an element is in the **number** of neutrons.

2.12. Recalling that isotopes of an element must all have the same number of protons, the following are isotopes of element X: $^{32}_{16}\text{X}$ and $^{36}_{16}\text{X}$

Key Experiments Developing Atomic Structure

2.13. The accepted mass of a proton is 1.672622×10^{-24} g, while that for the electron is 9.109384×10^{-28} g. The ratio of these two masses is:

$$\frac{\text{mass of proton}}{\text{mass of electron}} = \frac{1.672622 \times 10^{-24}\ \text{g}}{9.109384 \times 10^{-28}\ \text{g}} = 1836$$

So, Thomson's estimate was "off" by a factor of 2.

2.14. Negatively charged electrons in the cathode-ray tube collide with He atoms, splitting the atom into an electron and a He^+ cation.

2.15. One can observe from the accompanying graphic, that the β particles are attracted to the "+" plate, while the α particles are attracted to the "–" one. The gamma rays are attracted to neither.

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From this information, we know that α particles are **positively charged**, and that β particles are **negatively charged**. With the α particle being essentially the nucleus of a He atom, it is the heavier of the two particles.

2.16. Atoms are not solid, hard, or impenetrable. We know, thanks to the Kinetic-Molecular Theory, that the atoms are in rapid motion at all temperatures above absolute zero. They have mass (an important aspect of Dalton's hypothesis). The Rutherford "gold foil" experiment proved that atoms were *not* impenetrable, solid or hard, as alpha particles penetrated the atom and revealed a better picture of the atom.

Atomic Weight

2.17. Thallium has two stable isotopes ^{203}Tl and ^{205}Tl. The more abundant isotope is:___?___

The atomic weight of thallium is 204.4 u. The fact that this weight is closer to 205 than 203 indicates that the **205 isotope is the more abundant** isotope. Recall that the atomic weight is the "weighted average" of all the isotopes of each element. Hence the more abundant isotope will have a "greater contribution" to the atomic weight than the less abundant one.

2.18. Strontium has an atomic weight of 87.62, nearly 88, so ^{88}Sr is the most abundant.

2.19. The atomic mass of lithium is: (0.076)(6.015123) + (0.924)(7.016003) = 6.94

Recall that the atomic mass is a weighted average of all isotopes of an element, and is obtained by **adding** the *product* of (relative abundance × mass) for all isotopes.

2.20. (^{24}Mg mass)(% abundance/100) + (^{25}Mg mass)(% abundance/100) + (^{26}Mg mass)(% abundance/100) = atomic weight of Mg
(23.985042)(0.7899) + (24.985837)(0.1000) + (25.982593)(0.1101) = 24.31

2.21. The average atomic weight of gallium is 69.723 (from the periodic table). If we let $\boldsymbol{x}$ represent the abundance of the lighter isotope, and $(\mathbf{1} - \boldsymbol{x})$ the abundance of the heavier isotope, the expression to calculate the atomic weight of gallium may be written:

$(\boldsymbol{x})(68.9256) + (\mathbf{1} - \boldsymbol{x})(70.9247) = 69.723$

[Note that the sum of all the isotopic abundances must add to 100% -- or 1 (in decimal notation).]

Simplifying the equation gives:

$$68.9256\,\boldsymbol{x} + 70.9247 - 70.9247\,\boldsymbol{x} = 69.723$$
$$-1.9991\,\boldsymbol{x} = (69.723 - 70.9247)$$
$$-1.9992\,\boldsymbol{x} = -1.2017$$
$$\boldsymbol{x} = 0.6011$$

So, the relative abundance of gallium-69 is 60.11 % and that of gallium-71 is 39.89 %.

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2.22. Let x represent the abundance of ^{151}Eu and $(1 - x)$ represent the abundance of ^{153}Eu.
$(x)(150.9199) + (1 - x)(152.9212) = 151.96$, and solving for x yields
$x = 0.4803$; ^{151}Eu abundance is 48% (2 sf), ^{153}Eu abundance is 52%

The Periodic Table

2.23. Comparison of Titanium and Thallium:

Name	Symbol	Atomic #	Atomic Weight	Group #	Period #	Metal, Metalloid, or Nonmetal
Titanium	Ti	22	47.867	4B (4)	4	Metal
Thallium	Tl	81	204.38	3A (13)	6	Metal

2.24. Elements with symbols beginning with S:

	silicon	tin	antimony	sulfur	selenium
Symbol	Si	Sn	Sb	S	Se
At. number	14	50	51	16	34
Period	3	5	5	3	4
Group	4A (14)	4A (14)	5A (15)	6A (16)	6A (16)
	metalloid	metal	metalloid	nonmetal	nonmetal

2.25. Periods with 8 elements: **2**; Periods 2 (at. no. 3-10) and 3 (at. no. 11-18)

Periods with 18 elements: **2**; Periods 4 (at. no 19-36) and 5 (at. no. 37-54)

Periods with 32 elements: **2**; Periods 6 (at. no. 55-86) and 7 (at. no. 87-118)

2.26. There are 32 elements in the seventh period, many of them are called the Actinides, most are metals, and many are human-made elements.

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2.27. Elements fitting the following descriptions:

	Description	Elements
(a)	Nonmetals	C, Cl
(b)	Main group elements	C, Ca, Cl, Cs
(c)	Lanthanides	Ce
(d)	Transition elements	Cr, Co, Cd, Cu, Ce, Cf, Cm
(e)	Actinides	Cf, Cm
(f)	Gases	Cl

2.28. Name and chemical symbol for:

(a) C, carbon; N, nitrogen; O, oxygen; F, fluorine; Ne, neon
(b) Rb, rubidium

(c) Cl, chlorine

(d) H, hydrogen; He, helium; N, nitrogen; O, oxygen; F, fluorine; Ne, neon; Cl, chlorine; Ar, argon; Kr, krypton; Xe, xenon; Rn, radon are all gases at 20°C and 1 atmosphere pressure.

2.29. Classify the elements as metals, metalloids, or nonmetals:

	Metals	Metalloids	Nonmetals
N			X
Na	X		
Ni	X		
Ne			X
Np	X		

2.30. Match symbol with description:

(a) A radioactive element Bk
(b) A liquid at room temperature Br

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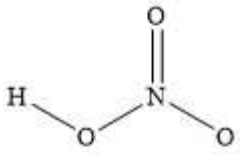
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(c) A metalloid	B
(d) An alkaline earth metal	Ba
(e) A group 5A (15) element	Bi

Molecules: Formulas, Models and Names

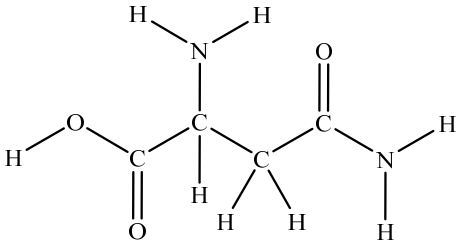
2.31. The molecular formula for nitric acid is HNO_3.

The structural formula is shown below (without lone-pair electrons). The structure around the N is flat. The oxygens are in a trigonal-planar arrangement, and the H–O–N atoms reside in bent arrangement.



Nitric acid

2.32. Molecular formula for asparagine (an amino acid) is $C_4H_8N_2O_3$. The structural formula is:



2.33. Names of binary molecular compounds

(a) NF_3	nitrogen trifluoride
(b) HI	hydrogen iodide
(c) BI_3	boron triiodide
(d) PF_5	phosphorus pentafluoride

2.34. Names of binary molecular compounds

(a) N_2O_5	dinitrogen pentaoxide
(b) P_4S_3	tetraphosphorus trisulfide
(c) OF_2	oxygen difluoride
(d) XeF_4	xenon tetrafluoride

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2.35. Formulas for:

(a) sulfur dichloride SCl_2

(b) dinitrogen pentaoxide N_2O_5

(c) silicon tetrachloride $SiCl_4$

(d) diboron trioxide B_2O_3

2.36. Formulas for:

(a) bromine trifluoride BrF_3

(b) xenon difluoride XeF_2

(c) hydrazine N_2H_4

(d) diphosphorus tetrafluoride P_2F_4

(e) butane C_4H_{10}

Ions and Ion Charges

2.37. Most commonly observed charge of monatomic ion for:

(a) Magnesium: 2+ —like all the alkaline earth metals

(b) Zinc: 2+

(c) Nickel: 2+

(d) Gallium: 3+ —an analog of aluminum

2.38. Most commonly observed charge of monatomic ion for:

(a) Sodium: 1+ —like all the alkali metals

(b) Aluminum: 3+

(c) Iron: 2+ and 3+ (transition metals frequently form multiple cations)

(d) Copper: 1+ and 2+ (transition metals frequently form multiple cations)

2.39. Most commonly observed charge of monatomic ion for:

(a) Selenium: 2–

(b) Fluorine: 1–

(c) Oxygen: 2–

(d) Nitrogen: 3–

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2.40. Most commonly observed charge of monatomic ion for:

(a) Bromine: 1–

(b) Sulfur: 2–

(c) Phosphorus: 3–

(d) Chlorine: 1–

2.41. The symbol and charge for the following ions:

(a)	barium	Ba^{2+}		(e)	sulfide	S^{2-}
(b)	titanium(IV)	Ti^{4+}		(f)	perchlorate	ClO_4^-
(c)	phosphate	PO_4^{3-}		(g)	cobalt(II)	Co^{2+}
(d)	hydrogen carbonate	HCO_3^-		(h)	sulfate	SO_4^{2-}

2.42. Symbol and charge for the ions:

(a)	permanganate	MnO_4^-		(d)	ammonium	NH_4^+
(b)	nitrite	NO_2^-		(e)	phosphate	PO_4^{3-}
(c)	dihydrogen phosphate	$H_2PO_4^-$		(f)	sulfite	SO_3^{2-}

2.43. When potassium becomes a monatomic ion, potassium—like all alkali metals—**loses 1 electron.** The noble gas atom with the same number of electrons as the potassium ion is **argon**.

2.44. Both atoms gain two electrons. O^{2-} has the same number of electrons as Ne and S^{2-} has the same number of electrons as Ar.

2.45. Number of protons and electrons in these ions.

Ion	**# of protons**	**# of electrons**
Na^+	11	10
Mg^{2+}	12	10
F^-	9	10
O^{2-}	8	10

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2.46. Number of protons and electrons in these ions.

Ion	# of protons	# of electrons
Ca^{2+}	20	18
Cu^{+}	29	28
Cu^{2+}	29	27
Cl^{-}	17	18

Ionic Compounds

2.47. Barium is in Group 2A (2) and is expected to form a 2+ ion while bromine is in group 7A (17) and expected to form a 1– ion. Since the compound would have to have an **equal amount** of negative and positive charges, the formula would be $BaBr_2$.

2.48. Cobalt is a transition metal. The Roman numeral indicates that this is a 3+ ion. Fluorine is in group 7A (17) and expected to form a 1– ion. Since the compound must have an **equal amount** of negative and positive charges, the formula would be CoF_3.

2.49. Formula, Charge, Number, and Name of ions in:

		cation & name	# of	anion & name	# of
(a)	K_2S	K^{+} potassium	2	S^{2-} sulfide	1
(b)	$CoSO_4$	Co^{2+} cobalt(II)	1	SO_4^{2-} sulfate	1
(c)	$KMnO_4$	K^{+} potassium	1	MnO_4^{-} permanganate	1
(d)	$(NH_4)_3PO_4$	NH_4^{+} ammonium	3	PO_4^{3-} phosphate	1
(e)	$Ca(ClO)_2$	Ca^{2+} calcium	1	ClO^{-} hypochlorite	2
(f)	$NaCH_3CO_2$	Na^{+} sodium	1	$CH_3CO_2^{-}$ acetate	1

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2.50. Formula, Charge, Number, and Name of ions in:

		cation & name	# of	anion & name	# of
(a)	$Mg(CH_3CO_2)_2$	Mg^{2+} magnesium	1	$CH_3CO_2^-$ acetate	2
(b)	$Al(OH)_3$	Al^{3+} aluminum	1	OH^- hydroxide	3
(c)	$CuCO_3$	Cu^{2+} copper(II)	1	CO_3^{2-} carbonate	1
(d)	$Ti(SO_4)_2$	Ti^{4+} titanium(IV)	1	SO_4^{2-} sulfate	2
(e)	KH_2PO_4	K^+ potassium	1	$H_2PO_4^-$ dihydrogen phosphate	1
(f)	$CaHPO_4$	Ca^{2+} calcium	1	HPO_4^{2-} hydrogen phosphate	1

2.51. Regarding cobalt oxides: cobalt(II) oxide, CoO cobalt ion : Co^{2+}

cobalt(III) oxide, Co_2O_3 cobalt ion : Co^{3+}

2.52. Platinum (II and IV) ions with chloride and sulfide ions:

(a) Pt^{2+}: $PtCl_2$ Pt^{4+}: $PtCl_4$
(b) Pt^{2+}: PtS Pt^{4+}: PtS_2

2.53. Provide correct formulas for compounds:

(a) $AlCl_2$ Incorrect; Correct is $AlCl_3$.

(b) KF_2 Incorrect; Correct is KF.

(c) Ga_2O_3 Correct.

(d) MgS Correct.

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2.54. Provide correct formulas for compounds if given formula is incorrect:

(a) Ca_2O Incorrect; Correct is CaO.

(b) $SrBr_2$ Correct as is.

(c) Fe_2O_5 Incorrect; Correct is either Fe_2O_3 for iron(III) oxide or FeO for iron(II) oxide

(d) Li_2O Correct as is.

2.55. Names for the ionic compounds

(a) K_2S potassium sulfide

(b) $CoSO_4$ cobalt(II) sulfate

(c) $(NH_4)_3PO_4$ ammonium phosphate

(d) $Ca(ClO)_2$ calcium hypochlorite

2.56. Name the following:

(a) $Ca(CH_3CO_2)_2$, calcium acetate

(b) $Ni_3(PO_4)_2$, nickel(II) phosphate

(c) $Al(OH)_3$, aluminum hydroxide

(d) KH_2PO_4, potassium dihydrogen phosphate

2.57. Names of the ionic compounds:

(a) Na_2S, sodium sulfide

(b) Na_2SO_3, sodium sulfite

(c) $NaHSO_4$, sodium hydrogen sulfate

(d) Na_2SO_4, sodium sulfate

2.58. Names of the ionic compounds:

(a) Li_3N, lithium nitride

(b) $LiNO_2$, lithium nitrite

(c) $LiNO_3$, lithium nitrate

2.59. Formulas for the ionic compounds

(a) ammonium carbonate $(NH_4)_2CO_3$

(b) calcium iodide CaI_2

(c) copper(II) bromide $CuBr_2$

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(d) aluminum phosphate $AlPO_4$

(e) silver(I) acetate $AgCH_3CO_2$

2.60. Formula for the following:

(a) calcium hydrogen carbonate $Ca(HCO_3)_2$

(b) potassium permanganate $KMnO_4$

(c) magnesium perchlorate $Mg(ClO_4)_2$

(d) potassium hydrogen phosphate K_2HPO_4

(e) sodium sulfite Na_2SO_3

2.61. Names and formulas for ionic compounds:

	anion	anion
cation	**CO_3^{2-}**	**I^-**
Na^+	Na_2CO_3 sodium carbonate	NaI sodium iodide
Ba^{2+}	$BaCO_3$ barium carbonate	BaI_2 barium iodide

2.62. Write formulas and name:

	anion	anion
cation	**NO_3^-**	**PO_4^{3-}**
Mg^{2+}	$Mg(NO_3)_2$ magnesium nitrate	$Mg_3(PO_4)_2$ magnesium phosphate
Fe^{3+}	$Fe(NO_3)_3$ iron(III) nitrate	$FePO_4$ iron(III) phosphate

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Coulomb's Law

2.63. The fluoride ion has a smaller radius than the iodide ion. Hence the distance between the sodium and fluoride ions will be less than the comparable distance between sodium and iodide. Coulomb's Law indicates that the attractive force becomes greater as the distance between the charges grows smaller—hence NaF will have stronger forces of attraction.

2.64. The attractive forces are stronger in CaO because the ion charges are greater (+2/–2 in CaO and +1/–1 in NaCl).

Atoms and the Mole

2.65. The mass, in grams of:

(a) $3.2 \text{ mol Al}\left(\frac{26.982 \text{ g Al}}{1 \text{ mol Al}}\right) = 86 \text{ g Al (2 sf)}$

(b) $2.35 \times 10^{-3} \text{ mol Fe}\left(\frac{55.845 \text{ g Fe}}{1 \text{ mol Fe}}\right) = 0.131 \text{ g Fe (3 sf)}$

(c) $0.12 \text{ mol Ca}\left(\frac{40.078 \text{ g Ca}}{1 \text{ mol Ca}}\right) = 4.8 \text{ g Ca (2 sf)}$

(d) $23.0 \text{ mol Ne}\left(\frac{20.180 \text{ g Ne}}{1 \text{ mol Ne}}\right) = 464 \text{ g Ne (3 sf)}$

Note that, whenever possible, one should use a molar mass of the substance that contains at least **one more** significant figure than the data, to reduce round-off error. Correct number of significant figures indicated by sf in parentheses.

2.66. The mass, in grams of:

(a) $1.24 \text{ mol Au}\left(\frac{196.97 \text{ g Au}}{1 \text{ mol Au}}\right) = 244 \text{ g Au (3 sf)}$

(b) $14.8 \text{ mol He}\left(\frac{4.0026 \text{ g He}}{1 \text{ mol He}}\right) = 59.2 \text{ g He (3 sf)}$

(c) $0.43 \text{ mol Pt}\left(\frac{195.08 \text{ g Pt}}{1 \text{ mol Pt}}\right) = 84 \text{ g Pt (2 sf)}$

(d) $2.42 \times 10^{-4} \text{ mol Rh}\left(\frac{102.91 \text{ g Rh}}{1 \text{ mol Rh}}\right) = 0.0249 \text{ g Rh (3 sf)}$

2.67. The amount (moles) of substance represented by:

(a) $87.21 \text{ g Cu}\left(\frac{1 \text{ mol Cu}}{63.546 \text{ g}}\right) = 1.372 \text{ mol Cu (4 sf)}$

(b) $0.024 \text{ g Na}\left(\frac{1 \text{ mol Na}}{22.990 \text{ g Na}}\right) = 0.0010 \text{ mol Na (2 sf)}$

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(c) $2.0\text{ mg Ir}\left(\frac{1\text{ g Ir}}{1000\text{ mg Ir}}\right)\left(\frac{1\text{ mol Ir}}{192.22\text{ g Ir}}\right) = 1.0\times10^{-5}\text{ mol Ir (2 sf)}$

(d) $6.75\text{ g Au}\left(\frac{1\text{ mol Au}}{196.97\text{ g Au}}\right) = 3.43\times10^{-2}\text{ mol Au (3 sf)}$

2.68. The amount (moles) of substance represented by:

(a) $9.4\text{ g Li}\left(\frac{1\text{ mol Li}}{6.94\text{ g Li}}\right) = 1.4\text{ mol Li (2 sf)}$

(b) $0.942\text{ g Sn}\left(\frac{1\text{ mol Sn}}{118.71\text{ g Sn}}\right) = 7.94\times10^{-3}\text{ mol Sn (3 sf)}$

(c) $0.037\text{ g Pt}\left(\frac{1\text{ mol Pt}}{195.08\text{ g Pt}}\right) = 1.9\times10^{-4}\text{ mol Pt (2 sf)}$

(d) $1.84\text{ g Xe}\left(\frac{1\text{ mol Xe}}{131.29\text{ g Xe}}\right) = 1.40\times10^{-2}\text{ mol Xe (3 sf)}$

2.69. The average mass of an atom (in grams) of each of these elements:

(a) He $\left(\frac{4.0026\text{ g He}}{1\text{ mol He}}\right)\left(\frac{1\text{ mol He}}{6.02214\times10^{23}\text{ atoms He}}\right) = 6.6465\times10^{-24}\text{ g/He atom (5 sf)}$

(b) F $\left(\frac{18.998\text{ g F}}{1\text{ mol F}}\right)\left(\frac{1\text{ mol F}}{6.02214\times10^{23}\text{ atoms F}}\right) = 3.1547\times10^{-23}\text{ g/F atom (5 sf)}$

(c) Sc $\left(\frac{44.956\text{ g Sc}}{1\text{ mol Sc}}\right)\left(\frac{1\text{ mol Sc}}{6.02214\times10^{23}\text{ atoms Sc}}\right) = 7.4651\times10^{-23}\text{ g/Sc atom (5 sf)}$

(d) Bi $\left(\frac{208.98\text{ g Bi}}{1\text{ mol Bi}}\right)\left(\frac{1\text{ mol Bi}}{6.02214\times10^{23}\text{ atoms Bi}}\right) = 3.4702\times10^{-22}\text{ g/Bi atom (5 sf)}$

2.70. The average mass of an atom (in grams) of each of these elements:

(a) Be $\left(\frac{9.0122\text{ g Be}}{1\text{ mol Be}}\right)\left(\frac{1\text{ mol Be}}{6.02214\times10^{23}\text{ atoms Be}}\right) = 1.4965\times10^{-23}\text{ g/Be atom (5 sf)}$

(b) Al $\left(\frac{26.982\text{ g Al}}{1\text{ mol Al}}\right)\left(\frac{1\text{ mol Al}}{6.02214\times10^{23}\text{ atoms Al}}\right) = 4.4805\times10^{-23}\text{ g/Al atom (5 sf)}$

(c) Kr $\left(\frac{83.798\text{ g Kr}}{1\text{ mol Kr}}\right)\left(\frac{1\text{ mol Kr}}{6.02214\times10^{23}\text{ atoms Kr}}\right) = 1.3915\times10^{-22}\text{ g/Kr atom (5 sf)}$

(d) Hg $\left(\frac{200.59\text{ g Hg}}{1\text{ mol Hg}}\right)\left(\frac{1\text{ mol Hg}}{6.02214\times10^{23}\text{ atoms Hg}}\right) = 3.3309\times10^{-22}\text{ g/Hg atom (5 sf)}$

2.71. 1.0-gram samples of He, Fe, Li, Si, C:

Which sample contains the **largest number** of atoms? …the **smallest number** of atoms?

If we calculate the number of atoms of any one of these elements, say He, the process is:

$1.0\text{ g He}\left(\frac{1\text{ mol He}}{4.0026\text{ g He}}\right)\left(\frac{6.022\times10^{23}\text{ He atoms}}{1\text{ mol He}}\right) = 1.5\times10^{23}\text{ He atoms (2 sf)}$

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All the calculations proceed analogously, with the ONLY numerical difference attributable to the molar mass of the element. Therefore, the element with the **smallest** molar mass (He) will have the **largest number** of atoms, while the element with the **largest** molar mass (Fe) will have the **smallest number** of atoms. This is a great question to answer by **thinking** rather than by calculating.

2.72. List samples in increasing amounts:

$$0.10 \text{ g K} \cdot \frac{1 \text{ mol K}}{39.098 \text{ g K}} = 0.0026 \text{ mol K}$$

$$0.10 \text{ g Mo} \cdot \frac{1 \text{ mol Mo}}{95.95 \text{ g Mo}} = 0.0010 \text{ mol Mo}$$

$$0.10 \text{ g Cr} \cdot \frac{1 \text{ mol Cr}}{51.996 \text{ g Cr}} = 0.0019 \text{ mol Cr}$$

$$0.10 \text{ g Al} \cdot \frac{1 \text{ mol Al}}{26.982 \text{ g Al}} = 0.0037 \text{ mol Al}$$

$0.0010 \text{ mol Mo} < 0.0019 \text{ mol Cr} < 0.0026 \text{ mol K} < 0.0037 \text{ mol Al}$

Because there are equal masses of each sample the element with the largest molar mass will have the smallest number of moles. The element with the smallest molar mass will have the largest number of moles.

2.73. Analysis of a 10.0-g sample of apatite contained 3.99g Ca, 1.85g P, 4.14g O, and 0.020g H. Calculating moles of each element gives:

$$3.99 \text{ g Ca} \cdot \frac{1 \text{ mol Ca}}{40.08 \text{ g Ca}} = 0.0996 \text{ mol Ca and } 1.85 \text{ g P} \cdot \frac{1 \text{ mol P}}{30.97 \text{ g P}} = 0.0597 \text{ mol P}$$

$$4.14 \text{ g O} \cdot \frac{1 \text{ mol O}}{16.00 \text{ g O}} = 0.259 \text{ mol O and } 0.020 \text{ g H} \cdot \frac{1 \text{ mol H}}{1.01 \text{ g H}} = 0.020 \text{ mol H}$$

From smallest to largest # moles: $0.020 \text{ mol H} < 0.0597 \text{ mol P} < 0.0996 \text{ mol Ca} < 0.259 \text{ mol O}$

Unlike SQ 2.72, each sample has a different mass. Consequently, we must calculate each one to determine the number of moles.

2.74. Semiconducting material contains 52 g Ga, 9.5 g Al, 112 g As. Element with largest number of atoms:

$$52 \text{ g Ga}\left(\frac{1 \text{ mol Ga}}{69.7 \text{ g Ga}}\right)\left(\frac{6.022 \times 10^{23} \text{ atoms Ga}}{1 \text{ mol Ga}}\right) = 4.5 \times 10^{23} \text{ atoms Ga}$$

$$9.5 \text{ g Al}\left(\frac{1 \text{ mol Al}}{27.0 \text{ g Al}}\right)\left(\frac{6.022 \times 10^{23} \text{ atoms Al}}{1 \text{ mol Al}}\right) = 2.1 \times 10^{23} \text{ atoms Al}$$

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$$112 \text{ g As}\left(\frac{1 \text{ mol As}}{74.92 \text{ g As}}\right)\left(\frac{6.022 \times 10^{23} \text{ atoms As}}{1 \text{ mol As}}\right) = 9.00 \times 10^{23} \text{ atoms As}$$

Arsenic has the largest number of atoms in the mixture.

Molecules, Compounds, and the Mole

2.75. Moles of carbon in 1.0 mol and 0.20 mol of glucose:

The molecular formula of glucose, $C_6H_{12}O_6$, tells us the number of moles of each element in one mole of the compound. In one mole of glucose, there are 6 mol C, 12 mol H, and 6 mol O. Thus, in 1.0 mol of glucose, there are 6.0 mol C.

$$\text{In 0.20 mol of glucose, there are } (0.20 \text{ mol glucose})\left(\frac{6 \text{ mol C}}{1 \text{ mol glucose}}\right) = 1.2 \text{ mol C}$$

2.76. Moles of hydrogen and oxygen in 1.0 mol and 4.0 mol of water:

The molecular formula of water, H_2O, tells us the number of moles of each element in one mole of the compound. In 1.0 mol of water, there are 2.0 mol H and 1.0 mol O.

$$\text{In 4.0 mol of water, there are } 4.0 \text{ mol water}\left(\frac{2 \text{ mol H}}{1 \text{ mol water}}\right) = 8.0 \text{ mol H}$$

$$\text{In 4.0 mol of water, there are } 4.0 \text{ mol water}\left(\frac{1 \text{ mol O}}{1 \text{ mol water}}\right) = 4.0 \text{ mol O}$$

2.77. Molar mass of the following: (with atomic weights expressed to 4 significant figures)

(a) Fe_2O_3 (2)(55.85) + (3)(16.00) = 159.70 g/mol

(b) BCl_3 (1)(10.81) + (3)(35.45) = 117.16 g/mol

(c) $C_6H_8O_6$ (6)(12.01) + (8)(1.008) + (6)(16.00) = 176.12 g/mol

(d) $Mg(NO_3)_2$ (1)(24.31) + (2)(14.01) + (6)(16.00) = 148.33 g/mol

2.78. Molar mass of the following: (with atomic weights expressed to 4 significant figures)

(a) $CaCO_3$ (1)(40.08) + 1(12.01) + 3(16.00) = 100.09 g/mol

(b) $Fe(C_6H_{11}O_7)_2$ (1)(55.85) + (12)(12.01) + (22)(1.008) + (14)(16.00) = 446.15 g/mol

(c) $CH_3CH_2CH_2CH_2SH$ (4)(12.01) + (10)(1.008) + (1)(32.06) = 90.18 g/mol

(d) $C_{20}H_{24}N_2O_2$ (20)(12.01) + (24)(1.008) + (2)(14.01) + (2)(16.00) = 324.41 g/mol

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2.79. Molar mass of the following: (with atomic weights expressed to 4 significant figures)

(a) $Ni(NO_3)_2 \cdot 6\ H_2O$ (1)(58.69) + (2)(14.01) + 6(16.00) + (12)(1.008) + (6)(16.00)
= 290.81 g/mol

(b) $CuSO_4 \cdot 5\ H_2O$ (1)(63.55) + (1)(32.06) + 4(16.00) + (10)(1.008) + (5)(16.00)
= 249.69 g/mol

2.80. Molar mass of the following:

(a) $H_2C_2O_4 \cdot 2\ H_2O$ (6)(1.008) + (2)(12.01) + (6)(16.00) = 126.07 g/mol
(b) $MgSO_4 \cdot 7\ H_2O$ (1)(24.31) + (1)(32.06) + (11)(16.00) + (14)(1.008) = 246.48 g/mol

2.81. Mass represented by 0.0255 moles of the following compounds:
Molar masses are calculated as before. To determine the mass represented by 0.0255 moles, recall that 1 mol of a substance has a mass equal to the molar mass expressed in units of grams.

We calculate for (a) the mass represented by 0.0255 moles:

$$0.0255 \text{ mol } C_3H_7OH\left(\frac{60.10 \text{ g } C_3H_7OH}{1 \text{ mol } C_3H_7OH}\right) = 1.53 \text{ g } C_3H_7OH$$

Compound	Molar mass	Mass of 0.0255 moles
(a) C_3H_7OH	60.10	1.53 g
(b) $C_{11}H_{16}O_2$	180.2	4.60 g
(c) $C_9H_8O_4$	180.2	4.60 g
(d) C_3H_6O	58.08	1.48 g

Masses are expressed to 3 sf, since the # of moles has 3.

2.82. Mass of 0.123 mol of each of the following:

(a) $0.123 \text{ mol } C_{14}H_{10}O_4\left(\frac{242.2 \text{ g } C_{14}H_{10}O_4}{1 \text{ mol } C_{14}H_{10}O_4}\right) = 29.8 \text{ g } C_{14}H_{10}O_4$

(b) $0.123 \text{ mol } C_4H_8N_2O_2\left(\frac{116.1 \text{ g } C_4H_8N_2O_2}{1 \text{ mol } C_4H_8N_2O_2}\right) = 14.3 \text{ g } C_4H_8N_2O_2$

(c) $0.123 \text{ mol } C_5H_{10}S\left(\frac{102.2 \text{ g } C_5H_{10}S}{1 \text{ mol } C_5H_{10}S}\right) = 12.6 \text{ g } C_5H_{10}S$

(d) $0.123 \text{ mol } C_{12}H_{17}NO\left(\frac{191.3 \text{ g } C_{12}H_{17}NO}{1 \text{ mol } C_{12}H_{17}NO}\right) = 23.5 \text{ g } C_{12}H_{17}NO$

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2.83. Regarding sulfur trioxide:

1. Amount of SO_3 in 1.00 kg: $1.00 \times 10^3 \text{ g } SO_3 \left(\frac{1 \text{ mol } SO_3}{80.06 \text{ g } SO_3} \right) = 12.5 \text{ mol } SO_3$

2. Number of SO_3 molecules:

$$12.5 \text{ mol } SO_3 \left(\frac{6.022 \times 10^{23} \text{ molecules } SO_3}{1 \text{ mol } SO_3} \right) = 7.52 \times 10^{24} \text{ molecules } SO_3$$

3. Number of S atoms: With 1 S atom per SO_3 molecule—

$$7.52 \times 10^{24} \text{ molecules } SO_3 \left(\frac{1 \text{ atom S}}{1 \text{ molecule } SO_3} \right) = 7.52 \times 10^{24} \text{ atoms S}$$

4. Number of O atoms: With 3 O atoms per SO_3 molecule—

$$7.52 \times 10^{24} \text{ molecules } SO_3 \left(\frac{3 \text{ atoms O}}{1 \text{ molecule } SO_3} \right) = 2.26 \times 10^{25} \text{ atoms O}$$

2.84. Number of ammonium and sulfate ions present:

$$0.20 \text{ mol } (NH_4)_2SO_4 \left(\frac{2 \text{ mol } NH_4^+}{1 \text{ mol } (NH_4)_2SO_4} \right) \left(\frac{6.022 \times 10^{23} \text{ } NH_4^+ \text{ ions}}{1 \text{ mol } NH_4^+} \right) = 2.4 \times 10^{23} \text{ } NH_4^+ \text{ ions}$$

$$0.20 \text{ mol } (NH_4)_2SO_4 \left(\frac{1 \text{ mol } SO_4^{2-}}{1 \text{ mol } (NH_4)_2SO_4} \right) \left(\frac{6.022 \times 10^{23} \text{ } SO_4^{2-} \text{ ions}}{1 \text{ mol } SO_4^{2-}} \right) = 1.2 \times 10^{23} \text{ } SO_4^{2-} \text{ ions}$$

$$0.20 \text{ mol } (NH_4)_2SO_4 \left(\frac{2 \text{ mol N}}{1 \text{ mol } (NH_4)_2SO_4} \right) \left(\frac{6.022 \times 10^{23} \text{ N atoms}}{1 \text{ mol N}} \right) = 2.4 \times 10^{23} \text{ N atoms}$$

$$0.20 \text{ mol } (NH_4)_2SO_4 \left(\frac{8 \text{ mol H}}{1 \text{ mol } (NH_4)_2SO_4} \right) \left(\frac{6.022 \times 10^{23} \text{ H atoms}}{1 \text{ mol H}} \right) = 9.6 \times 10^{23} \text{ H atoms}$$

$$0.20 \text{ mol } (NH_4)_2SO_4 \left(\frac{1 \text{ mol S}}{1 \text{ mol } (NH_4)_2SO_4} \right) \left(\frac{6.022 \times 10^{23} \text{ S atoms}}{1 \text{ mol S}} \right) = 1.2 \times 10^{23} \text{ S atoms}$$

$$0.20 \text{ mol } (NH_4)_2SO_4 \left(\frac{4 \text{ mol O}}{1 \text{ mol } (NH_4)_2SO_4} \right) \left(\frac{6.022 \times 10^{23} \text{ O atoms}}{1 \text{ mol O}} \right) = 4.8 \times 10^{23} \text{ O atoms}$$

2.85. $C_8H_9NO_2$, the molecular formula for acetaminophen has a molar mass of 151.17 g/mol.

$$2 \text{ caplets} \left(\frac{500 \text{ mg } C_8H_9NO_2}{1 \text{ caplet}} \right) \left(\frac{1 \text{ g}}{1000 \text{ mg}} \right) \left(\frac{1 \text{ mol } C_8H_9NO_2}{151.17 \text{ g } C_8H_9NO_2} \right) \left(\frac{6.022 \times 10^{23} \text{ molecules } C_8H_9NO_2}{1 \text{ mol } C_8H_9NO_2} \right)$$
$$= 4 \times 10^{21} \text{ molecules } C_8H_9NO_2 \text{ (1 sf)}$$

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2.86. In an Alka-Seltzer tablet:

(a) $325 \text{ mg } C_9H_8O_4\left(\frac{1\text{g}}{10^3 \text{ mg}}\right)\left(\frac{1 \text{ mol } C_9H_8O_4}{180.2 \text{ g } C_9H_8O_4}\right) = 1.80\times10^{-3} \text{ mol } C_9H_8O_4$

$1916 \text{ mg } NaHCO_3\left(\frac{1\text{g}}{10^3 \text{ mg}}\right)\left(\frac{1 \text{ mol } NaHCO_3}{84.006 \text{ g } NaHCO_3}\right) = 2.281\times10^{-3} \text{ mol } NaHCO_3$

$1000. \text{ mg } C_6H_8O_7\left(\frac{1\text{g}}{10^3 \text{ mg}}\right)\left(\frac{1 \text{ mol } C_6H_8O_7}{191.12 \text{ g } C_6H_8O_7}\right) = 5.232\times10^{-3} \text{ mol } C_6H_8O_7$

(b) $1.80\times10^{-3} \text{ mol } C_9H_8O_4\left(\frac{6.022\times10^{23} \text{ molecules } C_9H_8O_4}{1 \text{ mol } C_9H_8O_4}\right) = 1.09\times10^{21} \text{ molecules } C_9H_8O_4$

Percent Composition

2.87. Mass percent for: [4 significant figures]

(a) PbS: Molar mass = (1)(207.2) + (1)(32.06) = 239.3 g/mol

$\%Pb = \frac{207.2 \text{ g Pb}}{239.3 \text{ g PbS}} \times 100\% = 86.60\%$

%S = 100.00% – 86.60% = 13.40%

(b) C_3H_8: Molar mass = 44.097 g/mol

$\%C = \frac{36.03 \text{ g C}}{44.097 \text{ g } C_3H_8} \times 100\% = 81.71\%$

%H = 100.00% – 81.71% = 18.29%

(c) $C_{10}H_{14}O$: Molar mass = 150.221 g/mol

$\%C = \frac{120.11 \text{ g C}}{150.221 \text{ g } C_{10}H_{14}O} \times 100\% = 79.956\% = 79.96\% \text{ (4 sf)}$

$\%H = \frac{14.112 \text{ g H}}{150.221 \text{ g } C_{10}H_{14}O} \times 100\% = 9.394\%$

%O = 100.00% – (79.96% + 9.394%) = 10.65 %

2.88. Mass percent of elements in the compounds:

(a) $\frac{(8)(12.01) \text{ g C}}{166.18 \text{ } C_8H_{10}N_2O_2} \cdot 100\% = 57.82\% \text{ C}$ $\quad \frac{(10)(1.008) \text{ g H}}{166.18 \text{ } C_8H_{10}N_2O_2} \cdot 100\% = 6.066\% \text{ H}$

$\frac{(2)(14.01) \text{ g N}}{166.18 \text{ } C_8H_{10}N_2O_2} \cdot 100\% = 16.86\% \text{ N}$ $\quad \frac{(2)(16.00) \text{ g O}}{166.18 \text{ } C_8H_{10}N_2O_2} \cdot 100\% = 19.26\% \text{ O}$

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(b) $\dfrac{(10)(12.01)\text{ g C}}{156.26\text{ g }C_{10}H_{20}O} \cdot 100\% = 76.86\%\text{ C}$ $\qquad$ $\dfrac{(20)(1.008)\text{ g H}}{156.26\text{ g }C_{10}H_{20}O} \cdot 100\% = 12.90\%\text{ H}$

$\dfrac{(1)(16.00)\text{ g O}}{156.26\text{ g }C_{10}H_{20}O} \cdot 100\% = 10.24\%\text{ O}$

(c) $\dfrac{(1)(58.93)\text{ g Co}}{237.92\text{ g }CoCl_2 \cdot 6\,H_2O} \cdot 100\% = 24.77\%\text{ Co}$

$\dfrac{(2)(35.45)\text{ g Cl}}{237.92\text{ g }CoCl_2 \cdot 6\,H_2O} \cdot 100\% = 29.80\%\text{ Cl}$

$\dfrac{(12)(1.008)\text{ g H}}{237.92\text{ g }CoCl_2 \cdot 6\,H_2O} \cdot 100\% = 5.084\%\text{ H}$

$\dfrac{(6)(16.00)\text{ g O}}{237.92\text{ g }CoCl_2 \cdot 6\,H_2O} \cdot 100\% = 40.35\%\text{ O}$

2.89. Mass of CuS to provide 10.0 g of Cu:

To calculate the weight percent of Cu in CuS, we need the respective atomic weights:

Cu = 63.546; S = 32.06 adding CuS gives a molecular weight of 95.61

The % of Cu in CuS is then: $\left(\dfrac{63.546\text{ g Cu}}{95.61\text{ g CuS}}\right)\times 100 = 66.47\%$

Now with this fraction (inverted) calculate the mass of CuS that will provide 10.0 g of Cu:

$10.0\text{ g Cu}\left(\dfrac{95.61\text{ g CuS}}{63.546\text{ g Cu}}\right) = 15.0\text{ g CuS}$

2.90. Mass percent of Ti in ilmenite; Mass of ilmenite to obtain 750 g of Ti:

$\dfrac{47.87\text{ g Ti}}{151.71\text{ g }FeTiO_3} \cdot 100\% = 31.55\%\text{ Ti}$

$750\text{ g Ti} \cdot \dfrac{100\text{ g }FeTiO_3}{31.55\text{ g Ti}} = 2.4 \times 10^3\text{ g }FeTiO_3$

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Empirical and Molecular Formulas

2.91. The empirical formula ($C_2H_3O_2$) would have a molar mass of 59.04 g/mol.

Since the molar mass is 118.1 g/mol we can write

$$\frac{1 \text{ empirical formula}}{59.04 \text{ g succinic acid}} \cdot \frac{118.1 \text{ g succinic acid}}{1 \text{ mol succinic acid}} = \frac{2.0 \text{ empirical formulas}}{1 \text{ mol succinic acid}}$$

So, the molecular formula contains 2 empirical formulas ($2 \times C_2H_3O_2$) or $C_4H_6O_4$.

2.92. Molecular formula for a compound:

Empirical formula mass = 58.06 g/mol; $\frac{116.1 \text{ g/mol molecular formula}}{58.06 \text{ g/mol empirical formula}}$ = 2 empirical formulas/molecular formula. The molecular formula is $C_4H_8N_2O_2$.

2.93. Provide the empirical or molecular formula for the following, as requested:

	Empirical Formula	Molar Mass (g/mol)	Molecular Formula
(a)	CH	26.0	**C_2H_2**
(b)	CHO	116.1	**$C_4H_4O_4$**
(c)	**CH_2**	**112.2**	C_8H_{16}

Note that we can calculate the mass of an empirical formula by adding the respective atomic weights (13.0 for CH, for example). The molar mass (26.0 for part (a)) is twice that of an empirical formula, so the molecular formula is 2 × empirical formula (or **C_2H_2** in part (a)).

2.94. Provide the empirical or molecular formula for the following, as requested:

	Empirical Formula	Molar Mass (g/mol)	Molecular Formula
(a)	$C_2H_3O_3$	150.1	**$C_4H_6O_6$**
(b)	C_3H_8	44.1	**C_3H_8**
(c)	**B_2H_5**	**53.3**	B_4H_{10}

2.95. Calculate the empirical formula of acetylene by calculating the mole ratios of carbon and hydrogen in 100 g of the compound.

$$92.26 \text{ g C} \cdot \frac{1 \text{ mol C}}{12.011 \text{g C}} = 7.681 \text{ mol C} \quad \text{and } 7.74 \text{ g H} \cdot \frac{1 \text{ mol H}}{1.008 \text{ g H}} = 7.68 \text{ mol H}$$

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Calculate the mole ratio: $\frac{7.681 \text{ mol C}}{7.68 \text{ mol H}} = \frac{1 \text{ mol C}}{1 \text{ mol H}}$

The mole ratio indicates that there is 1 C atom for 1 H atom (1:1). The **empirical formula is then CH**. The formula mass is 13.02. Given that the molar mass of the compound is 26.02 g/mol, there are two formula units per molecular unit, hence the **molecular formula for acetylene is C_2H_2**.

2.96. Empirical formula of a B–H compound:

The compound is 88.5% B and 100% – 88.5% = 11.5% H. Assume 100.0 g of compound.

$$88.5 \text{ g B}\left(\frac{1 \text{ mol B}}{10.81 \text{ g B}}\right) = 8.19 \text{ mol B} \quad 11.5 \text{ g H}\left(\frac{1 \text{ mol H}}{1.008 \text{ g H}}\right) = 11.4 \text{ mol H}$$

$\frac{11.4 \text{ mol H}}{8.19 \text{ mol B}} = \frac{1.39 \text{ mol H}}{1 \text{ mol B}} = \frac{7/5 \text{ mol H}}{1 \text{ mol B}} = \frac{7 \text{ mol H}}{5 \text{ mol B}}$ So the empirical formula is B_5H_7.

2.97. Determine the empirical and molecular formulas of cumene:

Assume 100.0 g of compound.

The percentage composition of cumene is 89.94% C and 100.00% – 89.94% = 10.06%H. Calculate the ratio of mol C: mol H as done in SQ95.

$$89.94 \text{ g C} \cdot \frac{1 \text{ mol C}}{12.011 \text{g C}} = 7.488 \text{ mol C}$$

$$10.06 \text{ g H} \cdot \frac{1 \text{ mol H}}{1.008 \text{ g H}} = 9.980 \text{ mol H}$$

Calculating the mole ratio:

$\frac{9.980 \text{ mol H}}{7.488 \text{ mol C}} = \frac{1.333 \text{ mol H}}{1 \text{ mol C}}$ or a ratio of 3C : 4H

So, the empirical formula for cumene is C_3H_4, with a formula mass of 40.07.

If the molar mass is 120.2 g/mol, then dividing the "empirical formula mass" into the molar mass gives: 120.2/40.07 or 3 empirical formulas **per** molar mass. The **molecular formula** is then 3 × C_3H_4 or C_9H_{12}.

2.98. Empirical and Molecular formula for sulflower:

Assume 100.0 g of compound.

$$57.17 \text{ g S}\left(\frac{1 \text{ mol S}}{32.06 \text{ g S}}\right) = 1.783 \text{ mol S} \qquad 42.83 \text{ mol C}\left(\frac{1 \text{ mol C}}{12.011 \text{ g C}}\right) = 3.566 \text{ mol C}$$

$\frac{3.566 \text{ mol C}}{1.783 \text{ mol S}} = \frac{2.000 \text{ mol C}}{1 \text{ mol S}}$

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So the empirical formula is C_2S. The molar mass of a C_2S entity is 56.08 g/mol, so divide that into the molar mass:

$\frac{448.70\ \text{g/mol}}{56.08\ \text{g/mol}} = 8$. The molecular formula is $C_{16}S_8$

2.99. Empirical and Molecular formula for mandelic acid:

Assume 100.0 g of compound.

$$63.15\ \text{g C} \cdot \frac{1\ \text{mol C}}{12.011\ \text{g C}} = 5.258\ \text{mol C} \qquad 5.30\ \text{g H} \cdot \frac{1\ \text{mol H}}{1.008\ \text{g H}} = 5.26\ \text{mol H}$$

$$31.55\ \text{g O} \cdot \frac{1\ \text{mol O}}{15.999\ \text{g O}} = 1.972\ \text{mol O}$$

Notice that the amounts (moles) of C and H are the same. Dividing by the smallest amount, we calculate the mole ratio for C and O:

$$\frac{5.258\ \text{mol C}}{1.972\ \text{mol O}} = \frac{2.666\ \text{mol C}}{1\ \text{mol O}} \text{ or } \frac{2\ 2/3\ \text{mol C}}{1\ \text{mol O}} \text{ or } \frac{8/3\ \text{mol C}}{1\ \text{mol O}}$$

So 3 mol O combine with 8 mol C and 8 mol H and the empirical formula is $C_8H_8O_3$.

The formula mass of $C_8H_8O_3$ is 152.15. Given the data that the molar mass of the compound is 152.15 g/mL, the molecular formula for mandelic acid is $C_8H_8O_3$.

2.100. Empirical and Molecular formula for nicotine:

Assume 100.0 g of compound.

$$74.0\ \text{g C}\left(\frac{1\ \text{mol C}}{12.01\ \text{g C}}\right) = 6.16\ \text{mol C} \qquad 8.65\ \text{g H}\left(\frac{1\ \text{mol H}}{1.008\ \text{g H}}\right) = 8.58\ \text{mol H}$$

$$17.35\ \text{g N}\left(\frac{1\ \text{mol N}}{14.007\ \text{g N}}\right) = 1.239\ \text{mol N}$$

Now determine the mole:mole ratios:

$$\frac{6.16\ \text{mol C}}{1.239\ \text{mol N}} = \frac{4.97\ \text{mol C}}{1\ \text{mol N}} = \frac{5\ \text{mol C}}{1\ \text{mol N}} \qquad \frac{8.58\ \text{mol H}}{1.239\ \text{mol N}} = \frac{6.93\ \text{mol H}}{1\ \text{mol N}} = \frac{7\ \text{mol H}}{1\ \text{mol N}}$$

and the empirical formula is C_5H_7N.

Compare the mass of the empirical formula to that of the molecular formula: $\frac{116\ \text{g/mol}}{81.1\ \text{g/mol}} = 2$

so the molecular formula is $C_{10}H_{14}N_2$

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Determining Formulas from Mass Data

2.101. Given the masses of xenon involved, we can calculate the number of moles of the element:

$$0.674 \text{ g Xe} \cdot \frac{1 \text{mol Xe}}{131.29 \text{ g Xe}} = 0.00513 \text{ mol Xe}$$

The mass of fluorine present is: 0.869 g compound – 0.674 g Xe = 0.195 g F

$$0.195 \text{ g F} \cdot \frac{1 \text{ mol F}}{19.00 \text{ g F}} = 0.0103 \text{ mol F}$$

Calculating atomic ratios:

$\frac{0.0103 \text{ mol F}}{0.00513 \text{ mol Xe}} = \frac{2 \text{ mol F}}{1 \text{ mol Xe}}$ indicating that the empirical formula is XeF_2.

2.102. Value of x in the formula SF_x:

6.70 g compound – 1.47 g S = 5.23 g F; Determine mol of S and mol of F:

$$1.47 \text{ g S}\left(\frac{1 \text{ mol S}}{32.06 \text{ g S}}\right) = 0.0459 \text{ mol S} \qquad 5.23 \text{ g F}\left(\frac{1 \text{ mol F}}{18.998 \text{ g F}}\right) = 0.275 \text{ mol F}$$

$\frac{0.275 \text{ mol F}}{0.0459 \text{ mol S}} = \frac{6.00 \text{ mol F}}{1 \text{ mol S}}$, so the empirical formula is SF_6; x =

2.103. Value of x in the formula $BaCl_2 \cdot x\ H_2O$:

Molar mass of $BaCl_2$ = 208.23 g/mol

Molar mass of H_2O = 18.02 g/mol

Mass of H_2O evaporated by heating = 1.523 g – 1.298 g = 0.225 g H_2O evaporated

$$1.298 \text{ g } BaCl_2\left(\frac{1 \text{ mol } BaCl_2}{208.23 \text{ g } BaCl_2}\right) = 0.006233 \text{ mol } BaCl_2$$

$$0.225 \text{ g } H_2O\left(\frac{1 \text{ mol } H_2O}{18.02 \text{ g } H_2O}\right) = 0.0125 \text{ mol } H_2O$$

$$\text{Mole ratio of } H_2O \text{ to } BaCl_2 = \frac{0.0125 \text{ mol } H_2O}{0.006233 \text{ mol } BaCl_2} = \frac{2.00 \text{ mol } H_2O}{1 \text{ mol } BaCl_2}$$

Thus, there are 2 moles of H_2O for every 1 mole of $BaCl_2$ indicating that x has a value of 2 and the compound's formula is $BaCl_2 \cdot 2\ H_2O$.

The name of this compound is barium chloride dihydrate.

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2.104. Value of x in the formula $Na_2SO_4 \cdot x\ H_2O$:

Molar mass of Na_2SO_4 = 142.0 g/mol

Molar mass of H_2O = 18.02 g/mol

Mass of H_2O evaporated by heating = 2.343 g – 1.033 g = 1.310 g H_2O evaporated

$$1.033\ \text{g}\ Na_2SO_4\left(\frac{1\ \text{mol}\ Na_2SO_4}{142.0\ \text{g}\ Na_2SO_4}\right) = 0.007275\ \text{mol}\ Na_2SO_4$$

$$1.310\ \text{g}\ H_2O\left(\frac{1\ \text{mol}\ H_2O}{18.02\ \text{g}\ H_2O}\right) = 0.07270\ \text{mol}\ H_2O$$

$$\text{Mole ratio of } H_2O \text{ to } Na_2SO_4 = \frac{0.07270\ \text{mol}\ H_2O}{0.007275\ \text{mol}\ Na_2SO_4} = \frac{9.993\ \text{mol}\ H_2O}{1\ \text{mol}\ Na_2SO_4} \approx \frac{10\ \text{mol}\ H_2O}{1\ \text{mol}\ Na_2SO_4}$$

Thus, there are 10 moles of H_2O for every 1 mole of Na_2SO_4 indicating that x has a value of 10 and the compound's formula is $Na_2SO_4 \cdot 10\ H_2O$.

This compound is named sodium sulfate decahydrate.

2.105. Value of x in the formula $MgSO_4 \cdot x\ H_2O$:

Molar mass of $MgSO_4$ = 120.4 g/mol

Molar mass of H_2O = 18.02 g/mol

Molar mass of $MgSO_4 \cdot 7\ H_2O$ = 120.4 g + 7(18.02 g) = 246.5 g/mol $MgSO_4 \cdot 7\ H_2O$

In 1.394 g of $MgSO_4 \cdot 7\ H_2O$ there are

$$1.394\ \text{g}\ MgSO_4 \cdot 7\ H_2O\left(\frac{120.4\ \text{g}\ MgSO_4}{246.5\ \text{g}\ MgSO_4 \cdot 7\ H_2O}\right) = 0.6809\ \text{g}\ MgSO_4$$

In 0.885 g of $MgSO_4 \cdot x\ H_2O$ there are 0.6809 g $MgSO_4$ and 0.885 g – 0.6809 g = 0.204 g H_2O

$$0.6809\ \text{g}\ MgSO_4\left(\frac{1\ \text{mol}\ MgSO_4}{120.4\ \text{g}\ MgSO_4}\right) = 0.005655\ \text{mol}\ MgSO_4$$

$$0.204\ \text{g}\ H_2O\left(\frac{1\ \text{mol}\ H_2O}{18.02\ \text{g}\ H_2O}\right) = 0.0113\ \text{mol}\ H_2O$$

$$\text{Mole ratio of } H_2O \text{ to } MgSO_4 = \frac{0.0113\ \text{mol}\ H_2O}{0.005655\ \text{mol}\ MgSO_4} = \frac{2.00\ \text{mol}\ H_2O}{1\ \text{mol}\ MgSO_4}$$

Thus, in the partially hydrated sample there are 2 moles of H_2O for every 1 mole of $MgSO_4$ indicating that x has a value of 2 and the compound's formula is $MgSO_4 \cdot 2\ H_2O$.

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2.106. Formula for Ge_xCl_y:

3.69 g product – 1.25 g Ge = 2.44 g Cl; Determine # mol of Cl and Ge:

$$1.25 \text{ g Ge}\left(\frac{1 \text{ mol Ge}}{72.63 \text{ g Ge}}\right) = 0.0172 \text{ mol Ge} \qquad 2.44 \text{ g Cl}\left(\frac{1 \text{ mol Cl}}{35.45 \text{ g Cl}}\right) = 0.0688 \text{ mol Cl}$$

Calculate the mole ratio: $\frac{0.0688 \text{ mol Cl}}{0.0172 \text{ mol Ge}} = \frac{4.00 \text{ mol Cl}}{1 \text{ mol Ge}}$, so the empirical formula is $GeCl_4$.

Mass Spectrometry

2.107. Regarding the mass spectrum for NO_2:

(a) The four cations are: 14(N^+); 16(O^+); 30(NO^+); 46(NO_2^+). Note the assignments are based upon the masses of the atoms involved.

(b) The mass spectrum does provide evidence for NO bonds as opposed to OO bonds. An OO bond configuration would have produced a peak at 32 *m*/*Z*.

2.108. Regarding the mass spectrum for POF_3:

(a) Fragment at *m*/*Z* ratio of 85: The peak at 104 is the molecular ion, POF_3^+, so the fragment at 85 represents (104 – 19) or POF_2^+

(b) Fragment at *m*/Z ratio of 69: This fragment is PF_2^+, indicating a further loss of O (85 – 16).

(c) Fragment at *m*/*Z* ratio of 47 indicates a PO^+ fragment, and the fragment at *m*/*Z* ratio of 50 indicates a PF^+ fragment. The lack of a peak at 35, indicates an absence of OF^+.

2.109. Regarding the mass spectrum of CH_3Cl:

(a) Species responsible for lines at 50 and 52:

$^{12}CH_3{}^{35}Cl^+$ gives rise to the line at *m*/*Z* = 50 (12 + 3 + 35).

$^{12}CH_3{}^{37}Cl^+$ gives rise to the line at *m*/*Z* = 52 (12 + 3 + 37).

The differential in sizes of the lines is due to the relative distribution of isotopes of Cl, with the abundance of Cl-37 being about 1/3 that of Cl-35.

(b) The line at *m*/*Z* = 51 is due to $^{13}CH_3{}^{35}Cl^+$ with the accompanying smaller size of the peak. There is also a small contribution to this peak from $^{12}C^2H^1H_2{}^{35}Cl^+$.

2.110. Consider the peaks in the Br_2 mass spectrum at 158, 160, and 162:

(a) The *m*/*Z* peak at 158 is: $^{79}Br_2^+$; while the peak at *m*/*Z* 160 is: $^{79}Br^{81}Br^+$. Finally, the peak at *m*/*Z* 162 represents $^{81}Br_2^+$.

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(b) The abundances are close enough to assume an equal abundance of ^{79}Br and ^{81}Br. Two atoms from the two isotopes can be combined in four different manners to form Br_2: $^{79}Br_2$, $^{79}Br^{81}Br$, $^{81}Br^{79}Br$, and $^{81}Br_2$. Thus, the peak at *m/Z* 160 should have twice the intensity of the peaks at *m/Z* 158 and 162.

General Questions

2.111.

Symbol	^{58}Ni	^{33}S	^{20}Ne	^{55}Mn
Number of protons	28	16	10	25
Number of neutrons	30	17	10	30
Number of electrons in the neutral atom	28	16	10	25
Name of element	nickel	sulfur	neon	manganese

2.112. As the atomic weight of potassium is 39.098 (remember this is a weighted average), the lighter isotope, ^{39}K is more abundant than ^{41}K.

2.113. Crossword puzzle: Clues:

Horizontal

1–2 A metal used in ancient times: tin (Sn)

3–4 A metal that burns in air and is found in Group 5A(15): bismuth (Bi)

Vertical

1–3 A metalloid: antimony (Sb)

2–4 A metal used in U.S. coins: nickel (Ni)

Single squares:

1. A colorful nonmetal: sulfur (S)

2. A colorless gaseous nonmetal: nitrogen (N)

3. An element that makes fireworks green: boron (B)

4. An element that has medicinal uses: iodine (I)

Diagonal:

1–4 An element used in electronics: silicon (Si)

2–3 A metal used with Zr to make wires for superconducting magnets: niobium (Nb)

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Using these solutions, the following letters fit in the boxes

1 **S**	2 **N**
3 **B**	4 **I**

2.114. Regarding the abundance of elements:

(a) Mg is the most abundant main group metal.
(b) H is the most abundant nonmetal.
(c) Si is the most abundant metalloid.
(d) Fe is the most abundant transition element.
(e) F and Cl are the halogens included, and of these Cl is the most abundant.

2.115. Copper atoms:

(a) The average mass of one copper atom:

One mole of copper (with a mass of 63.546 g) contains 6.0221×10^{23} atoms. So, the average mass of **one** copper atom is: $\dfrac{63.546 \text{ g Cu/mol}}{6.02214 \times 10^{23} \text{ atoms Cu/mol}}$

$= 1.0552 \times 10^{-22}$ g/Cu atom

(b) Given the cost data: \$80.10 for 7.0 g and the mass of a Cu atom (from part (a)), the cost of one Cu atom is:

$$\frac{\$80.10}{7.0 \text{ g sample}}\left(\frac{100 \text{ g sample}}{99.999 \text{ g Cu}}\right)\left(\frac{1.0552 \times 10^{-22} \text{ g Cu}}{1 \text{ Cu atom}}\right) = 1.2 \times 10^{-21} \text{ dollars/Cu atom}$$

2.116. Identify which of the following is possible:

(a) silver foil that is 1.2×10^{-4} m thick (thick enough for an atom of Ag—possible)

(b) a sample of potassium that contains 1.784×10^{24} atoms (about 3 moles of K—possible)

(c) a gold coin of mass 1.23×10^{-3} kg (massive enough for more than one Au atom—possible)

(d) 3.43×10^{-27} mol S_8 (amount is less than one molecule of S_8—impossible)

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2.117. Identify the element that:

(a) Is in Group 2A (2) and the 5th period: Strontium

(b) Is in the 5th period and Group 4B (4): Zirconium

(c) Is in the second period in Group 4A (14): Carbon

(d) Is an element in the 4th period of Group 5A (15): Arsenic

(e) Is a halogen (Group 7A [17]) in the 5th period: Iodine

(f) Is an alkaline earth element (Group 2A [2]) in the 3rd period: Magnesium

(g) Is a noble gas (Group 8A [18]) in the 4th period: Krypton

(h) Is a nonmetal in Group 6A (16) and the 3rd period: Sulfur

(i) Is a metalloid in the 4th period: Germanium or Arsenic

2.118. **Carbon** has three allotropes. Graphite consists of flat sheets of carbon atoms, diamond has carbon atoms attached to four other others in a tetrahedron, and buckminsterfullerene is a 60-atom cage of carbon atoms. **Oxygen** has two allotropes. Diatomic oxygen consists of molecules containing two oxygen atoms and ozone consists of molecules containing three oxygen atoms. Other nonmetallic elements with allotropes include phosphorus (white, black and red), sulfur (more than 15 allotropic forms), and selenium (red, gray, and black).

2.119. Which of the following has the greater mass:

(a) 0.5 mol Na, 0.5 mol Si, 0.25 mol U

You can determine this without doing the entire calculation. Examine the molar mass of each element. 0.5 mol of any element has a mass that is one-half the molar mass. One quarter mol of U (molar mass approximately 238 g) will have the greatest mass of these three.

(b) 9.0 g of Na, 0.50 mol Na, 1.2×10^{22} atoms Na:

9.0 g Na is already in a unit of mass.

$$0.50 \text{ mol Na}\left(\frac{23.0 \text{ g Na}}{1 \text{ mol Na}}\right) = 12 \text{ g Na (2 sf)}$$

$$1.2 \times 10^{22} \text{ atoms Na}\left(\frac{1 \text{ mol Na}}{6.022 \times 10^{23} \text{ atoms Na}}\right)\left(\frac{23.0 \text{ g Na}}{1 \text{ mol Na}}\right) = 0.46 \text{ g Na}$$

0.50 mol Na will have the greatest mass of these three choices.

(c) 10 atoms of Fe or 10 atoms of K

As in (a), the full calculation is not needed. Fe has a greater atomic weight, so 10 atoms of Fe would have a greater mass than 10 atoms of K.

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2.120. Regarding the RDA of iron for women:

$$18 \text{ mg Fe}\left(\frac{1 \text{ g}}{1000 \text{ mg}}\right)\left(\frac{1 \text{ mol Fe}}{55.85 \text{ g Fe}}\right) = 3.2 \times 10^{-4} \text{ mol Fe}$$

$$3.2 \times 10^{-4} \text{ mol Fe}\left(\frac{6.02 \times 10^{23} \text{ atoms Fe}}{1 \text{ mol Fe}}\right) = 1.9 \times 10^{20} \text{ atoms Fe}$$

2.121. Arrange the elements from least massive to most massive:

Select a common unit by which to compare the substances (say grams?)

(a) 3.79×10^{24} atoms $\text{Fe}\left(\frac{1 \text{ mol Fe}}{6.022 \times 10^{23} \text{ atoms Fe}}\right)\left(\frac{55.85 \text{ g Fe}}{1 \text{ mol Fe}}\right) = 351 \text{ g Fe}$

(b) $19.92 \text{ mol } \text{H}_2\left(\frac{2.016 \text{ g } \text{H}_2}{1 \text{ mol } \text{H}_2}\right) = 40.16 \text{ g } \text{H}_2$

(c) $8.576 \text{ mol C}\left(\frac{12.011 \text{ g C}}{1 \text{ mol C}}\right) = 103.0 \text{ g C}$

(d) $7.40 \text{ mol Si}\left(\frac{28.085 \text{ g Si}}{1 \text{ mol Si}}\right) = 208 \text{ g Si}$

(e) $9.221 \text{ mol Na}\left(\frac{22.990 \text{ g Na}}{1 \text{ mol Na}}\right) = 212.0 \text{ g Na}$

(f) 4.07×10^{24} atoms $\text{Al}\left(\frac{1 \text{ mol Al}}{6.022 \times 10^{23} \text{ atoms Al}}\right)\left(\frac{26.982 \text{ g Al}}{1 \text{ mol Al}}\right) = 182 \text{ g Al}$

(g) $9.2 \text{ mol } \text{Cl}_2\left(\frac{70.90 \text{ g } \text{Cl}_2}{1 \text{ mol } \text{Cl}_2}\right) = 6.5 \times 10^2 \text{ g } \text{Cl}_2$

In ascending order of mass: H_2, C, Al, Si, Na, Fe, Cl_2

2.122. Regarding the atomic weights of P and O:

0.744 g phosphorus combined with (1.704 g – 0.744 g) = 0.960 g O

Expressing the mass per atom, divide P by 4 and O by 10: $\frac{(0.744/4) \text{ g P}}{(0.960/10) \text{ g O}} = \frac{1.94 \text{ g P}}{1 \text{ g O}}$

telling us that a P atom is 1.94 times more massive than a O atom. Using this ratio, we can determine the mass of a P atom if we define O as 16.00:

$$16.00 \text{ O}\left(\frac{1.94 \text{ g P}}{1 \text{ g O}}\right) = 31.0 \text{ P}$$

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2.123. Regarding the atomic weights of C and O:

0.876 g carbon combined with (3.210 g – 0.876 g) = 2.334 g O

Expressing the mass per atom, divide C by 1 and O by 2: $\frac{(2.334/2)\text{ g O}}{(0.876/1)\text{ g C}} = \frac{1.33\text{ g O}}{1\text{ g C}}$ telling us that an O atom is 1.33 times more massive than a C atom. Using this ratio, we can determine the mass of a O atom if we define C as 12.01: $12.01\text{ C}\left(\frac{1.33\text{ g O}}{1\text{ g C}}\right) = 16.0\text{ O}$

2.124. Percent of K in a Na-K alloy:

For every 100 atoms of the alloy, there are 68 atoms of K and 32 atoms of Na.

$$68\text{ atoms K}\left(\frac{1\text{ mol K}}{6.022 \times 10^{23}\text{ atoms K}}\right)\left(\frac{39.1\text{ g K}}{1\text{ mol K}}\right) = 4.4 \times 10^{-21}\text{ g K}$$

$$32\text{ atoms Na}\left(\frac{1\text{ mol Na}}{6.02 \times 10^{23}\text{ atoms Na}}\right)\left(\frac{23.0\text{ g Na}}{1\text{ mol Na}}\right) = 1.2 \times 10^{-21}\text{ g Na}$$

$$\text{mass \% K} = \frac{4.4 \times 10^{-21}\text{ g K}}{4.4 \times 10^{-21}\text{ g K} + 1.2 \times 10^{-21}\text{ g Na}} \times 100\% = 78\%\text{ K}$$

2.125. Possible compounds from ions:

	CO_3^{2-}	PO_4^{3-}
NH_4^+	$(NH_4)_2CO_3$	$(NH_4)_3PO_4$
Ni^{2+}	$NiCO_3$	$Ni_3(PO_4)_2$

Compounds are electrically neutral—hence the total positive charge contributed by the cations (+ ion) must equal the total negative charge contributed by the anions (– ion). Since carbonate is di-negative and phosphate is tri-negative, two ammonium ions are required for carbonate and three for phosphate. For the nickel(II) ion one carbonate ion is needed while three Ni(II) ions and two phosphate ions are necessary for neutral compounds.

2.126. A strontium atom has 38 electrons (same as atomic number). When an atom of strontium forms an ion, it loses two electrons, forming an ion having the same number of electrons as the noble gas krypton (with 36 electrons).

2.127. Compound from the list with the highest mass percent of Cl:

One way to answer this question is to calculate the %Cl in each of the five compounds. An observation that each compound has the same number of Cl atoms provides a "non-calculator" approach to answering the question.

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Since 3 Cl atoms will contribute the same TOTAL mass of Cl to the formula weights, the compound with the highest weight percent of Cl will also have the **lowest** weight percent of the other atom. Examining the atomic weights of the other atoms:

B	As	Ga	Al	P
10.81	74.92	69.72	26.98	30.97

B contributes the smallest mass of these five atoms, hence the smallest contribution to the molar masses of the five compounds—so BCl_3 has the highest weight percent of Cl.

2.128. Sample with largest number of ions:

These calculations require (1) converting mass to amount (moles) of compound, (2) converting amount of compound into amount of ions, and (3) converting amount of ions into numbers of ions. Each of the calculations is identical simply by substituting the appropriate molar mass of the compound and the number of ions formed per molecule. Note especially for (d) and (e), that the polyatomic ions remain intact—giving ONE ion PER polyatomic ion (e.g. CO_3^{2-} is 1 ion, not 4).

(a) $1.0\text{ g } CaCl_2\left(\frac{1\text{ mol } CaCl_2}{111\text{ g } CaCl_2}\right)\left(\frac{3\text{ mol ions}}{1\text{ mol } CaCl_2}\right)\left(\frac{6.02\times10^{23}\text{ ions}}{1\text{ mol ions}}\right) = 1.6\times10^{22}\text{ ions}$

(b) $1.0\text{ g } MgCl_2\left(\frac{1\text{ mol } MgCl_2}{95.2\text{ g } MgCl_2}\right)\left(\frac{3\text{ mol ions}}{1\text{ mol } MgCl_2}\right)\left(\frac{6.02\times10^{23}\text{ ions}}{1\text{ mol ions}}\right) = 1.9\times10^{22}\text{ ions}$

(c) $1.0\text{ g } CaS\left(\frac{1\text{ mol } CaS}{72.1\text{ g } CaS}\right)\left(\frac{2\text{ mol ions}}{1\text{ mol } CaS}\right)\left(\frac{6.02\times10^{23}\text{ ions}}{1\text{ mol ions}}\right) = 1.7\times10^{22}\text{ ions}$

(d) $1.0\text{ g } SrCO_3\left(\frac{1\text{ mol } SrCO_3}{148\text{ g } SrCO_3}\right)\left(\frac{2\text{ mol ions}}{1\text{ mol } SrCO_3}\right)\left(\frac{6.02\times10^{23}\text{ ions}}{1\text{ mol ions}}\right) = 8.1\times10^{21}\text{ ions}$

(e) $1.0\text{ g } BaSO_4\left(\frac{1\text{ mol } BaSO_4}{233\text{ g } BaSO_4}\right)\left(\frac{2\text{ mol ions}}{1\text{ mol } BaSO_4}\right)\left(\frac{6.02\times10^{23}\text{ ions}}{1\text{ mol ions}}\right) = 5.2\times10^{21}\text{ ions}$

$MgCl_2$ sample has the largest number of ions.

2.129. To determine the greater mass, let's first ask the question about the molar mass of adenine.

The formula for adenine is: $C_5H_5N_5$ with a molar mass of 135.13 g. The number of molecules requested is approximately 1/2 mole of adenine molecules. So, 1/2 mol of adenine molecules would have a mass of 1/2(135.13g) or 67.57 g. Thus, 3.0×10^{23} molecules of adenine has a greater mass than 40.0 g of adenine.

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2.130. Regarding BaF_2, $SiCl_4$, and $NiBr_2$:

(a) BaF_2: barium fluoride $SiCl_4$: silicon tetrachloride $NiBr_2$: nickel(II) bromide

(b) BaF_2 and $NiBr_2$ are ionic; $SiCl_4$ is molecular

(c) $0.50 \text{ mol } BaF_2 \left(\dfrac{175 \text{ g } BaF_2}{1 \text{ mol } BaF_2} \right) = 88 \text{ g } BaF_2$

$0.50 \text{ mol } SiCl_4 \left(\dfrac{170. \text{ g } SiCl_4}{1 \text{ mol } SiCl_4} \right) = 85 \text{ g } SiCl_4$

$1.0 \text{ mol } NiBr_2 \left(\dfrac{219 \text{ g } NiBr_2}{1 \text{ mol } NiBr_2} \right) = 2.2 \times 10^2 \text{ g } NiBr_2$

1.0 mol $NiBr_2$ has the largest mass

2.131. A drop of water has a volume of 0.050 mL. Assuming the density of water is 1.00 g/cm^3, the number of molecules of water may be calculated by first determining the mass of water present: $0.050 \text{ mL } H_2O \left(\dfrac{1 \text{ cm}^3}{1 \text{ mL}} \right) \left(\dfrac{1.00 \text{ g } H_2O}{1 \text{ cm}^3 \text{ } H_2O} \right) = 0.050 \text{ g } H_2O$

The molar mass of water is 18.02 g/mol.

$$0.050 \text{ g } H_2O \left(\frac{1 \text{ mol } H_2O}{18.02 \text{ g } H_2O} \right) \left(\frac{6.02 \times 10^{23} \text{ molecules } H_2O}{1 \text{ mol } H_2O} \right) = 1.7 \times 10^{21} \text{ molecules } H_2O$$

2.132. Regarding capsaicin:

(a) Molar mass = (18)(12.011) + (27)(1.008) + (1)(14.007) + (3)(15.999) = 305.418 g/mol

(b) $55 \text{ mg } C_{18}H_{27}NO_3 \left(\dfrac{1 \text{ g}}{10^3 \text{ mg}} \right) \left(\dfrac{1 \text{ mol } C_{18}H_{27}NO_3}{305.418 \text{ g } C_{18}H_{27}NO_3} \right) = 1.8 \times 10^{-4} \text{ mol } C_{18}H_{27}NO_3$

(c) $\dfrac{(18)(12.011) \text{ g C}}{305.418 \text{ g } C_{18}H_{27}NO_3} \cdot 100\% = 70.788\% \text{ C}$

$\dfrac{(27)(1.008) \text{ g H}}{305.418 \text{ g } C_{18}H_{27}NO_3} \cdot 100\% = 8.911\% \text{ H}$

$\dfrac{(1)(14.007) \text{ g N}}{305.418 \text{ g } C_{18}H_{27}NO_3} \cdot 100\% = 4.586\% \text{ N}$

$\dfrac{(3)(15.999) \text{ g O}}{305.418 \text{ g } C_{18}H_{27}NO_3} \cdot 100\% = 15.715\% \text{ O}$

(d) $55 \text{ mg } C_{18}H_{27}NO_3 \left(\dfrac{70.788 \text{ mg C}}{100 \text{ mg } C_{18}H_{27}NO_3} \right) = 39 \text{ mg C}$

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2.133. Molar mass and mass percent of the elements in $Cu(NH_3)_4SO_4 \cdot H_2O$:

Molar Mass: (1)(Cu) + (4)(N) + 12(H) + (1)(S) + (4)(O) + (2)(H) + (1)(O).
Combining the hydrogens and oxygen from water with the compound:

(1)(Cu) + (4)(N) + 14(H) + (1)(S) + (5)(O) =
(1)(63.546) + (4)(14.007) + 14(1.008) + (1)(32.06) + (5)(15.999) = 245.74 g/mol

The mass percentages are:

Cu: (63.546/245.74) × 100 = 25.86% Cu

N: (56.028/245.74) × 100 = 22.80% N

H: (14.112/245.74) × 100 = 5.743% H

S: (32.06/245.74) × 100 = 13.05% S

O: (79.995/245.74) × 100 = 32.55% O

The mass of copper and of water in 10.5 g of the compound:

$$\text{For Copper: } 10.5 \text{ g compound}\left(\frac{25.86 \text{ g Cu}}{100 \text{ g compound}}\right) = 2.72 \text{ g Cu}$$

$$\text{For Water: } 10.5 \text{ g compound}\left(\frac{18.02 \text{ g } H_2O}{245.74 \text{ g compound}}\right) = 0.770 \text{ g } H_2O$$

2.134. Molecular formula and molar mass for the compounds:

(a) Ethylene glycol $C_2H_6O_2$ Molar mass = 62.07 g/mol

$$\frac{(2)(12.01) \text{ g C}}{62.07 \text{ g } C_2H_6O_2} \cdot 100\% = 38.70\% \text{ C} \qquad \frac{(2)(16.00) \text{ g O}}{62.07 \text{ g } C_2H_6O_2} \cdot 100\% = 51.55\% \text{ O}$$

(b) Dihydroxyacetone $C_3H_6O_3$ Molar mass = 90.08 g/mol

$$\frac{(3)(12.01) \text{ g C}}{90.08 \text{ g } C_3H_6O_3} \cdot 100\% = 40.00\% \text{ C} \qquad \frac{(3)(16.00) \text{ g O}}{90.08 \text{ g } C_3H_6O_3} \cdot 100\% = 53.29\% \text{ O}$$

(c) Ascorbic acid $C_6H_8O_6$ Molar mass = 176.12 g/mol

$$\frac{(6)(12.01) \text{ g C}}{176.12 \text{ g } C_6H_8O_6} \cdot 100\% = 40.92\% \text{ C} \qquad \frac{(6)(16.00) \text{ g C}}{176.12 \text{ g } C_6H_8O_6} \cdot 100\% = 54.51\% \text{ O}$$

Ascorbic acid has the largest percentage of carbon **and** of oxygen.

2.135. The empirical formula of malic acid, if the ratio is: $C_1H_{1.50}O_{1.25}$

Since we prefer all subscripts to be integers, we ask what "multiplier" we can use to convert each of these subscripts to integers **while** retaining the given ratio of C:H:O. Multiplying each subscript by 4 (we need to convert the 0.25 to an integer) gives a ratio of $C_4H_6O_5$.

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2.136. Substance delivering greater number of Fe atoms:

$$100.\text{ mg } FeSO_4\left(\frac{1\text{ g}}{1\times10^3\text{ mg}}\right)\left(\frac{1\text{ mol } FeSO_4}{151.90\text{ g } FeSO_4}\right)\left(\frac{6.022\times10^{23}\text{ atoms Fe}}{1\text{ mol } FeSO_4}\right)=3.96\times10^{20}\text{ atoms Fe}$$

$$100.\text{ mg } Fe(C_6H_{11}O_7)_2\left(\frac{1\text{ g}}{1\times10^3\text{ mg}}\right)\left(\frac{1\text{ mol } Fe(C_6H_{11}O_7)_2}{446.14\text{ g } Fe(C_6H_{11}O_7)_2}\right)\left(\frac{6.022\times10^{23}\text{ atoms Fe}}{1\text{ mol } Fe(C_6H_{11}O_7)_2}\right)=1.35\times10^{20}\text{ atoms Fe}$$

The tablet containing $FeSO_4$ will deliver more atoms of iron.

2.137. A compound $Fe_x(CO)_y$ is 30.70% Fe: This implies that the balance of the mass (69.30%) is attributable to the CO molecules. One approach is to envision CO as **one** unit, with a molar mass of (12.011 + 15.999) or 28.010 g. Assuming we have 100 grams, the ratios of masses are then:

$$30.70\text{ g Fe}\left(\frac{1\text{ mol Fe}}{55.845\text{ g Fe}}\right)=0.5497\text{ mol Fe}$$

$$69.30\text{ g CO}\left(\frac{1\text{ mol CO}}{28.010\text{ g CO}}\right)=2.474\text{ mol CO}$$

Dividing by the smaller amount gives

$$\frac{2.474\text{ mol CO}}{0.5497\text{ mol Fe}}=\frac{4.501\text{ mol CO}}{1\text{ mol Fe}}=\frac{4\frac{1}{2}\text{ mol CO}}{1\text{ mol Fe}}=\frac{9\text{ mol CO}}{2\text{ mol Fe}}$$

The empirical formula is $Fe_2(CO)_9$.

2.138. Regarding ephedrine:

(a) $C_{10}H_{15}NO$ Molar mass = 165.236 g/mol

(b) $\dfrac{(10)(12.011)\text{ g C}}{165.236\text{ g } C_{10}H_{15}NO}\cdot 100\%=72.690\%\text{ C}$

(c) $0.125\text{ g } C_{10}H_{15}NO\left(\dfrac{1\text{ mol } C_{10}H_{15}NO}{165.236\text{ g } C_{10}H_{15}NO}\right)=7.56\times10^{-4}\text{ mol } C_{10}H_{15}NO$

(d)

$$7.56\times10^{-4}\text{ mol } C_{10}H_{15}NO\left(\frac{6.022\times10^{23}\text{ molecules } C_{10}H_{15}NO}{1\text{ mol } C_{10}H_{15}NO}\right)=4.56\times10^{20}\text{ molecules } C_{10}H_{15}NO$$

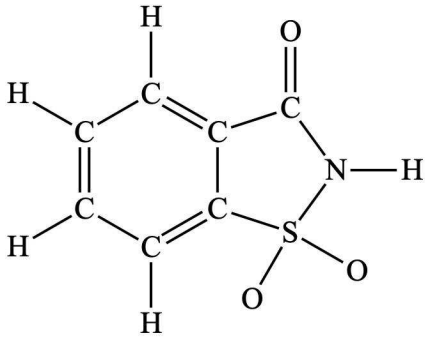
$$4.56\times10^{20}\text{ molecules } C_{10}H_{15}NO\left(\frac{10\text{ C atoms}}{1\text{ molecule } C_{10}H_{15}NO}\right)=4.56\times10^{21}\text{ C atoms}$$

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2.139. For the molecule saccharin, structural formula shown:

(a) The formula is $C_7H_5NO_3S$



saccharin

(b) Mol of saccharin associated with 125 mg:

$$125 \text{ mg saccharin}\left(\frac{1 \text{ g}}{1000 \text{ mg}}\right)\left(\frac{1 \text{ mol saccharin}}{183.18 \text{ g saccharin}}\right) = 6.82 \times 10^{-4} \text{ mol saccharin}$$

(c) Mass of S in 125 mg saccharin:

$$125 \text{ mg saccharin}\left(\frac{32.06 \text{ mg S}}{138.18 \text{ mg saccharin}}\right) = 21.9 \text{ mg S}$$

2.140. Name the compounds and indicate which are ionic:

Ionic compounds are in **bold** font.

(a) chlorine trifluoride
(b) nitrogen trichloride
(c) **strontium sulfate**
(d) **calcium nitrate**
(e) xenon tetrafluoride
(f) oxygen difluoride
(g) **potassium iodide**
(h) **aluminum sulfide**
(i) phosphorus trichloride
(j) **potassium phosphate**

2.141. Formulas for compounds; identify the ionic compounds

Ionic compounds are in **bold** font.

(a) sodium hypochlorite — **NaClO**
(b) boron triiodide — BI_3
(c) aluminum perchlorate — **$Al(ClO_4)_3$**
(d) calcium acetate — **$Ca(CH_3CO_2)_2$**
(e) potassium permanganate — **$KMnO_4$**

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(f) ammonium sulfite **$(NH_4)_2SO_3$**

(g) potassium dihydrogen phosphate **KH_2PO_4**

(h) disulfur dichloride S_2Cl_2

(i) chlorine trifluoride ClF_3

(j) phosphorus trifluoride PF_3

The ionic compounds are identified by looking for a *metal*.

2.142. Supply the indicated information:

Cation	Anion	Name	Formula
NH_4^+	Br^-	ammonium bromide	NH_4Br
Ba^{2+}	S^{2-}	barium sulfide	BaS
Fe^{2+}	Cl^-	iron(II) chloride	$FeCl_2$
Pb^{2+}	F^-	lead(II) fluoride	PbF_2
Al^{3+}	CO_3^{2-}	aluminum carbonate	$Al_2(CO_3)_3$
Fe^{3+}	O^{2-}	iron(III) oxide	Fe_2O_3

2.143. Empirical and molecular formulas:

(a) For fluorocarbonyl hypofluorite:

In a 100.00 g sample there are:

$$14.6 \text{ g C}\left(\frac{1 \text{ mol C}}{12.011 \text{ g C}}\right) = 1.22 \text{ mol C}$$

$$39.0 \text{ g O}\left(\frac{1 \text{ mol O}}{15.999 \text{ g O}}\right) = 2.44 \text{ mol O}$$

$$46.3 \text{ g F}\left(\frac{1 \text{ mol F}}{18.998 \text{ g F}}\right) = 2.44 \text{ mol F}$$

Dividing all three terms by 1.22 gives a ratio of O:C of 2:1. Likewise F:C is 2:1
The empirical formula would be CF_2O_2 with an "empirical mass" of 82.0 g/mol Since the molar mass is also 82.0 g/mol, the molecular formula is also CF_2O_2.

(b) For azulene:

Given the information that azulene is a hydrocarbon, if it is 93.71 % C, it is also (100.00 – 93.71) or 6.29 % H.

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In a 100.00 g sample of azulene there are

$$93.71 \text{ g C}\left(\frac{1 \text{ mol C}}{12.011 \text{ g C}}\right) = 7.802 \text{ mol C}$$

$$6.29 \text{ g H}\left(\frac{1 \text{ mol H}}{1.008 \text{ g H}}\right) = 6.24 \text{ mol H}$$

The ratio of C to H atoms is: 1.25 mol C : 1 mol H or a ratio of 5 mol C:4 mol H (C_5H_4). The mass of such an empirical formula is ~64. Given that the molar mass is ~128 g/mol, the molecular formula for azulene is $C_{10}H_8$.

2.144. Determine the empirical formula:

Assume 100.00 g of compound. Calculate moles of C, H, and As:

$$22.88 \text{ g C}\left(\frac{1 \text{ mol C}}{12.011 \text{ g C}}\right) = 1.905 \text{ mol C} \qquad 5.76 \text{ g H}\left(\frac{1 \text{ mol H}}{1.008 \text{ g H}}\right) = 5.71 \text{ mol H}$$

$$71.36 \text{ g As}\left(\frac{1 \text{ mol As}}{74.922 \text{ g As}}\right) = 0.9525 \text{ mol As}$$

Determine the mole ratios: $\frac{1.905 \text{ mol C}}{0.9525 \text{ mol As}} = \frac{2.000 \text{ mol C}}{1 \text{ mol As}}$ and

$\frac{5.71 \text{ mol H}}{0.9525 \text{ mol As}} = \frac{6.00 \text{ mol H}}{1 \text{ mol As}}$

The empirical formula is C_2H_6As. $\frac{210 \text{ g/mol}}{105.0 \text{ g/mol}} = 2.0$

With a molar mass of 210 g/mol, the molecular formula is $C_4H_{12}As_2$

2.145. Molecular formula of cadaverine:

Calculate the amount of each element in the compound (assuming you have 100 g)

$$58.77 \text{ g C}\left(\frac{1 \text{ mol C}}{12.011 \text{ g C}}\right) = 4.893 \text{ mol C}$$

$$13.81 \text{ g H}\left(\frac{1 \text{ mol H}}{1.008 \text{ g H}}\right) = 13.70 \text{ mol H}$$

$$27.40 \text{ g N}\left(\frac{1 \text{ mol N}}{14.007 \text{ g N}}\right) = 1.956 \text{ mol N}$$

The ratio of C:H:N can be found by dividing each by the smallest amount (1.956):
to give $C_{2.50}H_7N_1$ and converting each subscript to an integer (multiplying by 2)
$C_5H_{14}N_2$. The molar mass of this empirical formula would be approximately 102, hence the molecular formula is also $C_5H_{14}N_2$.

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2.146. For $Ni(CO)_x$, determine the value of x:

The mass of CO is: 0.364 g $Ni(CO)_x$ – 0.125 g Ni = 0.239 g CO

Determine moles of each:

$$0.239 \text{ g CO}\left(\frac{1 \text{ mol CO}}{28.01 \text{ g CO}}\right) = 0.00853 \text{ mol CO};\ 0.125 \text{ g Ni}\left(\frac{1 \text{ mol Ni}}{58.69 \text{ g Ni}}\right) = 0.00213 \text{ mol Ni}$$

The ratio of these is: $\frac{0.00853 \text{ mol CO}}{0.00213 \text{ mol Ni}} = \frac{4.01 \text{ mol CO}}{1 \text{ mol Ni}}$. The compound formula is $Ni(CO)_4$.

2.147. The empirical formula for MMT:

Moles of each atom present in 100. g of MMT:

$$49.5 \text{ g C}\left(\frac{1 \text{ mol C}}{12.011 \text{ g C}}\right) = 4.12 \text{ mol C}$$

$$3.2 \text{ g H}\left(\frac{1 \text{ mol H}}{1.008 \text{ g H}}\right) = 3.2 \text{ mol H}$$

$$22.0 \text{ g O}\left(\frac{1 \text{ mol O}}{15.999 \text{ g O}}\right) = 1.38 \text{ mol O}$$

$$25.2 \text{ g Mn}\left(\frac{1 \text{ mol Mn}}{54.938 \text{ g Mn}}\right) = 0.459 \text{ mol Mn}$$

The ratio of C:H:Mn:O can be found by dividing each by the smallest amount (0.459):
to give $C_9H_7MnO_3$.

2.148. Mass % of P; Mass of calcium phosphate to produce 15.0 kg P:

Molar mass of $Ca_3(PO_4)_2$ = 310.17 g/mol

Mass % of P in $Ca_3(PO_4)_2$: $\frac{2(30.97) \text{ g P}}{310.17 \text{ g } Ca_3(PO_4)_2} \times 100\% = 19.97\% \text{ P}$

With this percentage, we calculate the amount of the phosphate to produce the requested amount of P: $15.0 \text{ kg P} \cdot \frac{100 \text{ kg } Ca_3(PO_4)_3}{19.97 \text{ kg P}} = 75.1 \text{ kg } Ca_3(PO_4)_2$

2.149. Chromium(III) oxide has the formula Cr_2O_3.

The weight percent of Cr in Cr_2O_3 is: $\frac{(2 \cdot 52.00)}{(2 \cdot 52.00) + (3 \cdot 16.00)} \cdot 100\% = 68.42\%$ Cr.

[The numerator is the sum of the mass of 2 atoms of Cr, while the denominator is the sum of the mass of 2 atoms of Cr and 3 atoms of O.]

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The weight of Cr_2O_3 necessary to produce 850 kg Cr:

$$850 \text{ kg Cr}\left(\frac{100 \text{ kg } Cr_2O_3}{68.42 \text{ kg Cr}}\right) = 1200 \text{ kg } Cr_2O_3 \text{ (2 sf)}$$

The second fraction represents the %Cr in the oxide.

2.150. Mass percent of Sb in the sulfide; Mass of stibnite in 1.00 kg of ore:

$$\text{Percent of Sb: } \frac{(2)(121.76) \text{ g Sb}}{339.7 \text{ g } Sb_2S_3} \times 100\% = 71.69\% \text{ Sb}$$

$$\text{Mass of Stibnite in ore: } 1.00 \text{ kg ore}\left(\frac{1000 \text{ g}}{1 \text{ kg}}\right)\left(\frac{10.6 \text{ g Sb}}{100 \text{ g ore}}\right)\left(\frac{100 \text{ g } Sb_2S_3}{71.69 \text{ g Sb}}\right) = 148 \text{ g } Sb_2S_3$$

2.151. Empirical and Molecular formula for I_xCl_y:

I_2	+	Cl_2	→	I_xCl_y
0.678 g		(1.246 – 0.678) = 0.568 g		1.246 g

Calculate the ratio of I:Cl atoms

$$0.678 \text{ g I} \cdot \frac{1 \text{ mol I}}{126.9 \text{ g I}} = 5.34 \times 10^{-3} \text{ mol I atoms}$$

$$0.568 \text{ g Cl} \cdot \frac{1 \text{ mol Cl}}{35.45 \text{ g Cl}} = 1.60 \times 10^{-2} \text{ mol Cl atoms}$$

$$\text{The ratio of Cl:I is: } \frac{1.60 \times 10^{-2} \text{ mol Cl atoms}}{5.34 \times 10^{-3} \text{ mol I atoms}} = \frac{3.00 \text{ mol Cl atoms}}{1 \text{ mol I atoms}}$$

The empirical formula is ICl_3 (FW = 233.3)
Given that the molar mass of I_xCl_y was 467 g/mol, we can calculate the number of empirical formulas per mole:

$$\frac{467 \text{g/mol}}{233.3 \text{g/empiricalformula}} = 2 \frac{\text{empirical formulas}}{\text{mol}} \text{ for a molecular formula of } I_2Cl_6.$$

2.152. Empirical formula of the vanadium and sulfur compound:

Calculate the moles of each element:

$$2.04 \text{ g V}\left(\frac{1 \text{ mol V}}{50.94 \text{ g V}}\right) = 0.0400 \text{ mol V}$$

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$$1.93 \text{ g S}\left(\frac{1 \text{ mol S}}{32.06 \text{ g S}}\right) = 0.0602 \text{ mol S}$$

Now we can calculate the ratios of V:S

$$\frac{0.0602 \text{ mol S}}{0.0400 \text{ mol V}} = \frac{1.50 \text{ mol S}}{1 \text{ mol V}} = \frac{3 \text{ mol S}}{2 \text{ mol V}}$$

The empirical formula is V_2S_3.

2.153. Mass of Fe in 15.8 kg of FeS_2:

$$\% \text{ Fe in } FeS_2 = \frac{55.85 \text{ g Fe}}{119.97 \text{ g } FeS_2} \times 100\% = 46.55 \text{ \% Fe}$$

and in 15.8 kg FeS_2: $15.8 \text{ kg } FeS_2 \cdot \frac{46.55 \text{ kg Fe}}{100.00 \text{ kg } FeS_2} = 7.36 \text{ kg Fe}$

2.154. Statements about octane that are not true:

(a) True. $0.500 \text{ mol } C_8H_{18}\left(\frac{114.2 \text{ g } C_8H_{18}}{1 \text{ mol } C_8H_{18}}\right) = 57.1 \text{ g } C_8H_{18}$

(b) True. $\frac{(8)(12.01) \text{ g C}}{114.2 \text{ g } C_8H_{18}} \cdot 100\% = 84.1\% \text{ C}$

(c) True. Dividing the molecular formula by 2 gives the simplest integral ratio of C and H.

(d) False. $57.1 \text{ g } C_8H_{18} \cdot \frac{(18)(1.008) \text{ g H}}{114.2 \text{ g } C_8H_{18}} = 9.07 \text{ g H}$

2.155. The formula of barium molybdate is $BaMoO_4$. What is the formula for sodium molybdate? This question is answered by observing that the compound indicates ONE barium ion. Since the barium ion has a 2+ charge, 2 Na^+ cations would be needed, making the formula for sodium molybdate Na_2MoO_4 or choice (d).

2.156. Identity of the metal that forms MCl_4: The compound is 74.75% Cl, and we know that this mass represents 4 Cl atoms, so we determine the molar mass of the compound:

$$\frac{74.75 \text{ g Cl}}{100 \text{ g } MCl_4} = \frac{(4)(35.45) \text{ g Cl}}{\text{molar mass } MCl_4}$$; Solving for molar mass: 189.7 g/mol.

Determine the mass of the metal by difference: 189.7 g MCl_4 – (4)(35.45) g Cl = 47.9 g.

M is Ti, titanium.

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2.157. Mass of Bi in two tablets of Pepto-Bismol™ ($C_{21}H_{15}Bi_3O_{12}$):

Moles of the active ingredient:

$$2 \text{ tablets}\left(\frac{262 \text{ mg } C_{21}H_{15}Bi_3O_{12}}{1 \text{ tablet}}\right)\left(\frac{1 \text{ g}}{1000 \text{ mg}}\right)\left(\frac{1 \text{ mol } C_{21}H_{15}Bi_3O_{12}}{1086.3 \text{ } C_{21}H_{15}Bi_3O_{12}}\right) = 4.82 \times 10^{-4} \text{ mol } C_{21}H_{15}Bi_3O_{12}$$

Mass of Bi:

$$4.82 \times 10^{-4} \text{ mol } C_{21}H_{15}Bi_3O_{12}\left(\frac{3 \text{ mol Bi}}{1 \text{ mol } C_{21}H_{15}Bi_3O_{12}}\right)\left(\frac{208.98 \text{ g Bi}}{1 \text{ mol Bi}}\right) = 0.302 \text{ g Bi}$$

2.158. Molar mass of MO_2; Possible elements for M:

We know that the compound is 15.2% O, and the molecule has 2 O atoms. Calculate the molar mass: $\frac{15.2 \text{ g O}}{100 \text{ g } MO_2} = \frac{(2)(16.00) \text{ g O}}{\text{molar mass } MO_2}$. The molar mass for MO_2 = 211 g/mol.

Determine the mass of the metal by difference: 211 g MO_2 – (2)(16.00) g O = 179 g.
An examination of the periodic table lets us know that M is Hf, hafnium.

2.159. What is the molar mass of ECl_4 and the identity of E?

2.50 mol of ECl_4 has a mass of 385 grams. The molar mass of ECl_4 would be:
$\frac{385 \text{ g } ECl_4}{2.50 \text{ mol } ECl_4}$ = 154 g/mol ECl_4.

Since the molar mass is 154, and we know that there are 4 chlorine atoms per mole of the compound, we can subtract the mass of 4 chlorine atoms to determine the mass of E.
154 – 4(35.5) = 12. The element with an atomic mass of 12 g/mol is **carbon**.

2.160. Atomic weights of A and Z:

First, calculate the molar masses of A_2Z_3 and AZ_2 using the mass (g) and amount (moles) of each:

$$\frac{22.48 \text{ g } A_2Z_3}{0.15 \text{ mol } A_2Z_3} = 150 \text{ g/mol } A_2Z_3$$

$$\frac{12.44 \text{ g } AZ_2}{0.15 \text{ mol } AZ_2} = 83 \text{ g/mol } AZ_2$$

For AZ_2: (atomic weight A) + (2)(atomic weight Z) = 83; and

For A_2Z_3: (2)(atomic weight A) + (3)(atomic weight Z) = 150

Rearrange the AZ_2 equation to solve for (atomic weight A):
atomic weight A = 83 – 2(atomic weight Z)

Substitute this into the equation for A_2Z_3: (2)[83 – (2)(atomic weight Z)] + (3)(atomic weight Z) = 150.

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The atomic weight of Z = 16 g/mol and substituting that value into either of the two original equations gives an atomic weight A = 51 g/mol. These atomic weight values indicate that A is the element oxygen and Z is vanadium.

2.161. For what value of *n*, will Br compose 0.105% of the mass of the polymer, $Br_3C_6H_3(C_8H_8)_n$?

Knowing that the 3Br atoms comprises 0.105% of the formula weight of the polymer, we can write the fraction:

$\frac{3 \cdot \text{Br}}{\text{formula weight}} = 0.00105$ (where $3 \cdot \text{Br}$) is 3 times the atomic weight of Br. Substituting we get $\frac{239.7}{\text{formula weight}} = 0.00105$ or $\frac{239.7}{0.00105} = \text{formula weight} = 2.283 \times 10^5$ g.

Noting that the "fixed" part of the formula contains 3 Br atoms, 6 C atoms, and 3 H atoms, we can calculate the mass associated with this part of the molecule (3(79.904) + 6(12.011) + 3(1.008) = 314.802). Subtracting this mass from the total (2.283×10^5) gives 2.280×10^5 g as the mass corresponding to the "C_8H_8" units. Since each such unit has a mass of 104.15 (8C + 8H), we can divide the mass of *one* C_8H_8 unit into the 2.280×10^5 mass remaining:

$\frac{2.280 \times 10^5 \text{g}}{104.15 \text{g}/\text{C}_8\text{H}_8} = 2189\ \text{C}_8\text{H}_8 \text{ units}$ to give a value for *n* of 2.19×10^3 (3 sf).

2.162. The molar mass of hemoglobin represents 100% of the mass of the molecule, while Fe represents 0.335% of the molecule. Set up a ratio of mass of Fe: mass of hemoglobin, which parallels this percentage of Fe in the molecule:

$\frac{(4)(55.85)\text{ g Fe}}{\text{molar mass hemoglobin}} = \frac{0.335\text{ g Fe}}{100\text{ g hemoglobin}}$ and solving for (molar mass hemoglobin) gives 6.67×10^4 g/mol

2.163. For the ^{64}Zn atom:

(a) Calculate the density of the nucleus:

Assume the nucleus of the Zn atom is a sphere (whose volume is $4/3\ \pi r^3$.) Given the desired units, express the radius in units of cm and calculate nuclear volume:

$$4.8 \times 10^{-6}\ \text{nm}\left(\frac{100\ \text{cm}}{1 \times 10^9\ \text{nm}}\right) = 4.8 \times 10^{-13}\ \text{cm}$$

$$V \text{ of nucleus} = \frac{4}{3}\pi\left(4.8 \times 10^{-13}\ \text{cm}\right)^3 = 4.6 \times 10^{-37}\ \text{cm}^3$$

Then determine nuclear density

$$\text{density of nucleus} = \frac{1.06 \times 10^{-22}\ \text{g}}{4.6 \times 10^{-37}\ \text{cm}^3} = 2.3 \times 10^{14}\ \text{g/cm}^3$$

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(b) The density of the space occupied by the electrons: Express the radius of the atom in cm, as in part (a): 0.125 nm = 1.25×10^{-8} cm

Calculate the volume of that sphere: $V = \frac{4}{3}\pi\left(1.25 \times 10^{-8}\ \text{cm}\right)^3 = 8.18 \times 10^{-24}\ \text{cm}^3$

Mass of 30 electrons (the atomic number of zinc is 30): 30 electrons × 9.11×10^{-28} g = 2.73×10^{-26} g Density of space occupied by electrons =

$$\frac{2.73 \times 10^{-26}\ \text{g}}{8.18 \times 10^{-24}\ \text{cm}^3} = 3.34 \times 10^{-3}\ \text{g/cm}^3$$

(c) As we've learned, the mass of the atom is concentrated in the nucleus—borne out by these densities.

2.164. Estimate the radius of a Pb atom:

(a) Volume of cube = $(1.000\ \text{cm})^3 = 1.000\ \text{cm}^3$

$$1.000\ \text{cm}^3\ \text{Pb}\left(\frac{11.35\ \text{g Pb}}{1\ \text{cm}^3}\right)\left(\frac{1\ \text{mol Pb}}{207.2\ \text{g Pb}}\right)\left(\frac{6.022 \times 10^{23}\ \text{atoms Pb}}{1\ \text{mol Pb}}\right) = 3.299 \times 10^{22}\ \text{atoms Pb}$$

(b) Volume of one lead atom = $\frac{0.60 \times 1.000\ \text{cm}^3}{3.299 \times 10^{22}\ \text{atoms Pb}} = 1.8 \times 10^{-23}\ \text{cm}^3$

$1.8 \times 10^{-23}\ \text{cm}^3 = (^4/_3)(\pi)(\text{Pb radius})^3$ and solving, Pb radius = 1.6×10^{-8} cm

2.165. Calculate:

(a) moles of nickel—found by density **once** the volume of foil is calculated.

0.550 mm = 0.0550 cm

$V = 1.25\ \text{cm} \times 1.25\ \text{cm} \times 0.0550\ \text{cm} = 8.59 \times 10^{-2}\ \text{cm}^3$

$$\text{Mass} = 8.59 \times 10^{-2}\ \text{cm}^3\ \text{Ni}\left(\frac{8.902\ \text{g Ni}}{1\ \text{cm}^3\ \text{Ni}}\right) = 0.765\ \text{g Ni}$$

$$0.765\ \text{g Ni} \cdot \frac{1\ \text{mol Ni}}{58.69\ \text{g Ni}} = 1.30 \times 10^{-2}\ \text{mol Ni}$$

(b) Formula for the fluoride salt:

Mass F = (1.261 g salt – 0.765 g Ni) = 0.496 g F

Moles F = $0.496\ \text{g F} \cdot \frac{1\ \text{mol F}}{19.00\ \text{g F}} = 2.61 \times 10^{-2}$ mol F, so 1.30×10^{-2} mol Ni combines with 2.61×10^{-2} mol F, indicating a formula of NiF_2

(c) Name: nickel(II) fluoride

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2.166. Concerning uranium:

(a) 0.199 g U_xO_y – 0.169 g U = 0.030 g O

$$0.169 \text{ g U}\left(\frac{1 \text{ mol U}}{238.03 \text{ g U}}\right) = 7.10 \times 10^{-4} \text{ mol U}$$

$$0.030 \text{ g O}\left(\frac{1 \text{ mol O}}{16.0 \text{ g O}}\right) = 1.9 \times 10^{-3} \text{ mol O}$$

$$\frac{1.9 \times 10^{-3} \text{ mol O}}{7.10 \times 10^{-4} \text{ mol U}} = \frac{2.64 \text{ mol O}}{1 \text{ mol U}} \approx \frac{2\frac{2}{3} \text{ mol O}}{1 \text{ mol U}} = \frac{8 \text{ mol O}}{3 \text{ mol U}}$$

so the empirical formula is U_3O_8.

$$7.10 \times 10^{-4} \text{ mol U}\left(\frac{1 \text{ mol } U_3O_8}{3 \text{ mol U}}\right) = 2.37 \times 10^{-4} \text{ mol } U_3O_8$$

(b) The atomic weight of U is 238.03 u, implying that the isotope ^{238}U is the most abundant.

(c) 0.865 g – 0.679 g = 0.186 g H_2O lost upon heating

$$0.186 \text{ g } H_2O\left(\frac{1 \text{ mol } H_2O}{18.02 \text{ g } H_2O}\right) = 0.0103 \text{ mol } H_2O$$

$$0.679 \text{ g } UO_2(NO_3)_2\left(\frac{1 \text{ mol } UO_2(NO_3)_2}{394.0 \text{ g } UO_2(NO_3)_2}\right) = 0.00172 \text{ mol } UO_2(NO_3)_2$$

$$\frac{0.0103 \text{ mol } H_2O}{0.00172 \text{ mol } UO_2(NO_3)_2} = \frac{5.99 \text{ mol } H_2O}{1 \text{ mol } UO_2(NO_3)_2}$$

The formula of the hydrated compound is $UO_2(NO_3)_2 \cdot 6\ H_2O$.

2.167. The volume of a cube of Na containing 0.125 mol Na:

First, we need to know the **mass** of 0.125 mol Na:

$$0.125\text{mol Na} \cdot \frac{22.99 \text{ g Na}}{1 \text{ mol Na}} = 2.87 \text{ g Na} \quad (3 \text{ sf})$$

Now we can calculate the volume that contains 2.87 g Na: (using the density given)

$$2.87\text{g Na} \cdot \frac{1 \text{ cm}^3}{0.97 \text{ g Na}} = 3.0 \text{ cm}^3$$

To determine the length of one edge of a perfect cube we take the cube root of the volume.

$$\sqrt[3]{3.0 \text{ cm}^3} = 1.4 \text{ cm}$$

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2.168. Calculate the atomic weight:

Convert Abundances to percentages first, for example 0.185% = 0.00185.

Determine the contribution of each isotope by multiplying the abundance × isotopic mass:

(0.00185)(135.9071) + (0.00251)(137.9060) + (0.8845)(139.9054) + (0.1111)(141.9093) = 140.1 which is the atomic weight of the element cerium, Ce.

In the Laboratory

2.169. Molecules of water per formula unit of $MgSO_4$:

From 1.687 g of the hydrate, only 0.824 g of the magnesium sulfate remain.

The mass of water contained in the solid is: (1.687 – 0.824) or 0.863 grams
Use the molar masses of the solid and water to calculate the number of moles of each substance present:

$$0.824 \text{ g } MgSO_4\left(\frac{1 \text{ mol } MgSO_4}{120.36 \text{ g } MgSO_4}\right) = 6.85 \times 10^{-3} \text{ mol } MgSO_4$$

$$0.863 \text{ g } H_2O\left(\frac{1 \text{ mol } H_2O}{18.02 \text{ g } H_2O}\right) = 4.79 \times 10^{-2} \text{ mol } H_2O$$

The ratio of H_2O to $MgSO_4$ is: $\frac{4.79 \times 10^{-2} \text{ mol } H_2O}{6.85 \times 10^{-3} \text{ mol } MgSO_4} = \frac{7.00 \text{ mol } H_2O}{1 \text{ mol } MgSO_4}$

So we write the formula $MgSO_4 \cdot 7\ H_2O$.

2.170. Number of waters of hydration in alum:

Moles of anhydrous alum: 4.74 g hydrated compound – 2.16 g H_2O = 2.58 g $KAl(SO_4)_2$

$$2.58 \text{ g } KAl(SO_4)_2\left(\frac{1 \text{ mol } KAl(SO_4)_2}{258.2 \text{ g } KAl(SO_4)_2}\right) = 0.00999 \text{ mol } KAl(SO_4)_2$$

Moles of water: $2.16 \text{ g } H_2O\left(\frac{1 \text{ mol } H_2O}{18.02 \text{ g } H_2O}\right) = 0.120 \text{ mol } H_2O$. Now we determine the number of moles of water per mol of alum: $\frac{0.120 \text{ mol } H_2O}{0.00999 \text{ mol } KAl(SO_4)_2} = \frac{12.0 \text{ mol } H_2O}{1 \text{ mol } KAl(SO_4)_2}$

There are 12 water molecules per formula unit of $KAl(SO_4)_2$; $x = 12$

2.171. We can calculate the mass of Sn used in the compound by the **difference** of masses in the original mixture and the mass recovered after reaction: = 1.056 g – 0.601 g = 0.455 g Sn

The # mol of Sn is: $0.455 \text{ g Sn}\left(\frac{1 \text{ mol Sn}}{118.71 \text{ g Sn}}\right) = 0.00383 \text{ mol Sn}$

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Similarly for iodine, the mass of iodine can be converted into moles of **atomic iodine.**

$$1.947 \text{ g I}\left(\frac{1 \text{ mol I}}{126.90 \text{ g I}}\right) = 0.01534 \text{ mol I}$$

The empirical formula is determined by calculating the RATIO of **iodine** to **tin**.

$$\frac{0.01534 \text{ mol I}}{0.00383 \text{ mol Sn}} = \frac{4.00 \text{ mol I}}{\text{mol Sn}}$$

The empirical formula for the compound: Sn_1I_4, or more commonly expressed SnI_4.

2.172. Determine the empirical formula:

Assume 100.0 g of sample.

$$54.0 \text{ g C}\left(\frac{1 \text{ mol C}}{12.01 \text{ g}}\right) = 4.50 \text{ mol C} \qquad 6.00 \text{ g H}\left(\frac{1 \text{ mol H}}{1.008 \text{ g H}}\right) = 5.95 \text{ mol H}$$

$$40.0 \text{ g O}\left(\frac{1 \text{ mol O}}{16.00 \text{ g O}}\right) = 2.50 \text{ mol O}$$

Determine the mol ratio for C:O $\dfrac{4.50 \text{ mol C}}{2.50 \text{ mol O}} = \dfrac{1.80 \text{ mol C}}{1 \text{ mol O}} = \dfrac{1\frac{4}{5} \text{ mol C}}{1 \text{ mol O}} = \dfrac{9 \text{ mol C}}{5 \text{ mol O}}$

Determine the mol ratio for H:O $\dfrac{5.95 \text{ mol H}}{2.50 \text{ mol O}} = \dfrac{2.38 \text{ mol H}}{1 \text{ mol O}} \approx \dfrac{2\frac{2}{5} \text{ mol H}}{1 \text{ mol O}} = \dfrac{12 \text{ mol H}}{5 \text{ mol O}}$

Answer (d) $C_9H_{12}O_5$ is correct. Several reasons exist for an incorrect empirical formula. The students apparently did not correctly calculate the number of moles of material in 100.0 g or they improperly calculated the ratio of those moles in determining their empirical formula.

2.173. Using the student data, calculate the number of moles of $CaCl_2$ and moles of H_2O:

$$0.739 \text{ g } CaCl_2 \cdot \frac{1 \text{ mol } CaCl_2}{111.0 \text{ g } CaCl_2} = 0.00666 \text{ mol } CaCl_2$$

$$(0.832 \text{ g} - 0.739 \text{ g}) \text{ or } 0.093 \text{ g } H_2O \cdot \frac{1 \text{ mol } H_2O}{18.02 \text{ g } H_2O} = 0.0052 \text{ mol } H_2O$$

The number of moles of water/mol of calcium chloride is $\dfrac{0.0052 \text{ mol } H_2O}{0.00666 \text{ mol } CaCl_2} = 0.78$

This is a strong indication that they should **(c) heat the crucible again, and then reweigh it**.

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2.174. Empirical formula of tin oxide: 14.710 g crucible & Sn – 13.457 g crucible = 1.253 g Sn

$$1.253 \text{ g Sn}\left(\frac{1 \text{ mol Sn}}{118.71 \text{ g Sn}}\right) = 0.01056 \text{ mol Sn}$$

The O present is:15.048 g (crucible & Sn & O) – 14.710 g (crucible & Sn) = 0.338 g O

$$0.338 \text{ g O}\left(\frac{1 \text{ mol O}}{15.999 \text{ g O}}\right) = 0.0211 \text{ mol O}$$

Now the ratio of O/Sn is: 0.0211 mol O/0.01056 mol Sn = 2.00 mol O/1 mol Sn

The formula is SnO_2.

Summary and Conceptual Questions

2.175. Necessary information to calculate the number of atoms in one cm^3 of iron:

A sample calculation to arrive at an exact value is shown below:

$$1.00 \text{ cm}^3\left(\frac{7.87 \text{ g Fe}}{1 \text{ cm}^3}\right)\left(\frac{1 \text{ mol Fe}}{55.85 \text{ g Fe}}\right)\left(\frac{6.022 \times 10^{23} \text{ atoms Fe}}{1 \text{ mol Fe}}\right) = 8.49 \times 10^{22} \text{ atoms Fe}$$

Note that we needed: (d) density of iron, (b) molar mass of Fe, and (c) Avogadro's number.

2.176. Relationship between abundance and atomic number:

Element abundance generally decreases with increasing atomic number (with exceptions at Li–B and Sc–Fe). Elements with an even atomic number appear to be slightly more abundant than those with an odd atomic number.

2.177. Reactivity of Mg, Ca, and Ba:

(a) Given the greater reactivity of Ca over Mg with water, one would anticipate that Ba would be even more reactive than Ca or Mg—with a more vigorous release of hydrogen gas.

(b) Mg is in period 3, Ca in period 4, and Ba is in period 6. Reactivity of these metals increases down the group. (This trend is noted for Group 1A (1) as well.)

2.178. One possible method involves the following steps:

(1) weigh a representative sample of jelly beans (say 10 or so) in order to determine the average mass of a jelly bean;

(2) weigh the jelly beans in the jar (subtract the mass of the empty jar from the mass of the jar filled with jelly beans);

(3) use the average mass per jelly bean and the total mass of the jelly beans in the jar to determine the approximate number of jelly beans in the jar.

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Solution and Answer Guide

Kotz Treichel Townsend Treichel, Chemistry and Chemical Reactivity 11e, 978-0-357-85140-1, Chapter 3: Chemical Reactions

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Applying Chemical Principles

Superconductors

3.1.1. What is the value of x in a sample of $La_{2-x}Ba_xCuO_4$?

Given: %La = 63.43, %Ba = 5.085, %Cu = 15.69, and %O = 15.80.

Noting that the percentages supplied total to 100%, we can calculate the moles of each element present, assuming we have 100.g of the substance:

$$\frac{63.43\text{g La}}{1}\cdot\frac{1\text{mol La}}{138.91\text{g La}} = 0.4566\text{ mol La};\quad \frac{5.085\text{g Ba}}{1}\cdot\frac{1\text{mol Ba}}{137.33\text{g Ba}} = 0.03703\text{ mol Ba}$$

$$\frac{15.69\text{g Cu}}{1}\cdot\frac{1\text{mol Cu}}{63.546\text{g Cu}} = 0.2469\text{ mol Cu};\quad \frac{15.80\text{g O}}{1}\cdot\frac{1\text{mol O}}{15.999\text{g O}} = 0.9876\text{mol O}$$

Now unlike our usual procedure of dividing each number of moles by the **smallest**, we can exercise a bit of reasoning here. Note that the formula of the superconductor shows the subscript of Cu as 1. So let's force our calculations to have 1 has the subscript by dividing each of the amounts of La, Ba, and O by the moles of Cu:

$$\frac{0.4566\text{ mol La}}{0.2469\text{ mol Cu}} = 1.849\text{ mol La/mol Cu}\qquad \frac{0.03703\text{ mol Ba}}{0.2469\text{ mol Cu}} = 0.1500\text{ mol Ba/mol Cu}$$

$$\frac{0.9875\text{ mol O}}{0.2469\text{ mol Cu}} = 4.000\text{ mol O/mol Cu}$$

Note that this ratio simply **confirms** our formula for the superconductor. Since Ba is *x*, and we found $x = 0.1500$, then the subscript for La is (2-*x*) or 1.85. The formula for the semiconductor is then $La_{1.85}Ba_{0.15}CuO_4$.

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3.1.2. Percent by mass of the elements in $YBa_2Cu_3O_{6.93}$ if $x = 0.07$:

Determine mass of each of the elements:

$$1.000 \text{ mol Y} \cdot \frac{88.906 \text{g Y}}{1 \text{ mol Y}} = 88.906 \text{ g Y} \quad 2.000 \text{ mol Ba} \cdot \frac{137.33 \text{ g Ba}}{1 \text{ mol Ba}} = 274.7 \text{ g Ba}$$

$$3.000 \text{ mol Cu} \cdot \frac{63.546 \text{ g Cu}}{1 \text{ mol Cu}} = 190.6 \text{ g Cu} \qquad 6.930 \text{ mol O} \cdot \frac{15.999 \text{ g O}}{1 \text{ mol O}} = 110.9 \text{ g O}$$

The sum of these masses = 665.1 g. Now determine the % of each element:

$$\frac{88.906 \text{ g Y}}{665.1 \text{ g}} \cdot 100\ \% = 13.4\% \text{ Y} \qquad \frac{274.7 \text{ g Ba}}{665.1 \text{ g}} \cdot 100\ \% = 41.3\ \% \text{ Ba}$$

$$\frac{190.6 \text{ g Cu}}{665.1 \text{ g}} \cdot 100\ \% = 28.7\% \text{ Cu} \qquad \frac{110.9 \text{g O}}{665.1 \text{ g}} \cdot 100\ \% = 16.7\ \% \text{ O}$$

3.1.3. Charges present on the copper ions in $YBa_2Cu_3O_7$?

Since all the charges must add to 0, and since we're confident that each O will have a charge of 2−, we can add the charges for Y, Ba, and O to obtain (3+) + 2(2+) + 7(2−) = 7−. The three Cu ions must have charges that add to 7+. So, if we have TWO Cu ions at 2+ and ONE Cu ion at 3+, we accomplish that requirement.

3.1.4. A balanced chemical equation for the requested reaction:

$Y_2O_3(s) + 4\ BaCO_3(s) + 6\ CuO(s) \rightarrow 2\ YBa_2Cu_3O_{6.5}(s) + 4\ CO_2(g)$

Given the amount of Cu per mol of the product compound, the value of x is 0.5.

3.1.5. What mass of oxygen is required to convert 1.00 g $YBa_2Cu_3O_{6.50}$ to $YBa_2Cu_3O_{6.93}$?

Molar mass of YBCO: 1(Y) + 2(Ba) + 3(Cu) + 6.50(O)

= 1(88.906) + 2(137.33) + 3(63.546) + 6.50(15.999) = 658.20 g/mol

$$\frac{1.00 \text{ g YBCO}}{1} \cdot \frac{1 \text{ mol YBCO}}{658.20 \text{ g YBCO}} = 0.00152 \text{ mol YBCO}.$$

Adding (6.93 − 6.50) or 0.43 mol O_2 requires:

$$\frac{32.00 \text{ g O}_2}{1 \text{mol O}_2} \cdot \frac{1 \text{ mol O}_2}{2 \text{mol O}} \cdot \frac{0.43 \text{ mol O}}{1 \text{ mol YBCO}} \cdot \frac{0.00152 \text{ mol YBCO}}{1} = 0.010 \text{ g O}_2$$

Sequestering Carbon Dioxide

3.2.1. Balanced net ionic equation for the reaction of Ca^{2+} with H_2CO_3:

$Ca^{2+}(aq) + H_2CO_3(aq) + 2\ H_2O(\ell) \rightarrow CaCO_3(s) + 2\ H_3O^+(aq)$

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3.2.2. Protons, electrons, and neutrons in a C-14 atom:

Protons = 6; Electrons = 6 (it is neutral); Neutrons = (14 − 6) = 8

Black Smokers and Volcanoes

3.3.1. Formulas and names of ions making up compounds:

$CaSO_4$ Calcium ion, Ca^{2+}; sulfate ion, SO_4^{2-}

MnS Manganese(II) ion, Mn^{2+}; sulfide ion, S^{2-}

FeS Iron(II) ion, Fe^{2+}; sulfide ion, S^{2-}

NiS Nickel(II) ion, Ni^{2+}; sulfide ion, S^{2-}

3.3.2. Oxidation number for S in sulfide and sulfate ions:

S^{2-}: oxidation number = charge on ion for a monatomic ion, so −2

SO_4^{2-}: (S) + 4(O) = −2 (polyatomic ions have oxidation numbers that sum to charge on ions)
so (S) + 4(−2) = −2. A little algebra yields +6.

Practicing Skills

Introduction to Chemical Equations

3.1. $P_4(s) + 5\ O_2(g) \rightarrow P_4O_{10}(s)$

Reactants: $P_4(s)$, $O_2(g)$ Products: $P_4O_{10}(s)$

Stoichiometric coefficients: 1, 5, 1 (note that the 1's are implicit, and not typically written)

Designations s and g refer to the physical states of solid and gas respectively.

3.2. Equation from the description: $4\ NH_3(g) + 7\ O_2(g) \rightarrow 4\ NO_2(g) + 6\ H_2O(\ell)$

3.3. For the reaction: $P_4(s) + 6\ Cl_2(g) \rightarrow 4\ PCl_3(\ell)$ stoichiometric coefficients for the reactants are in a 1:6 ratio, so 8000 molecules of P_4 will require 48000 molecules of Cl_2.

3.4. For the reaction: $2\ Al(s) + 3\ Br_2(\ell) \rightarrow Al_2Br_6(s)$, stoichiometric coefficients for the reactants are in the ratio 2 Al: 3 Br_2. This ratio indicates that 6.0×10^{23} molecules of Br_2 will require 2/3 the number of Al atoms, or $\frac{2\text{ atoms Al}}{3\text{ molecules Br}_2} \cdot 6.0 \times 10^{23}\text{molecules Br}_2 = 4.0 \times 10^{23}\text{atoms Al}$

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3.5. Balanced equation for the oxidation of CO: $2\ CO(g) + O_2(g) \rightarrow 2\ CO_2(g)$
What mass of oxygen is required to convert 1.00 g of CO to 1.57 g CO_2? Law of Conservation of Matter indicates that 0.57 g of O_2 are needed to convert 1.00 g to 1.57 g.

3.6. The oxidation of Mg can be represented by the equation: $2\ Mg(s) + O_2(g) \rightarrow 2\ MgO(s)$.
If 0.20 mol of Mg form 0.20 mol of MgO, it will require HALF that amount of O_2, or 0.10 mol.

Balancing Equations

Balancing equations can be a matter of "running in circles" if a reasonable methodology is not employed. While there isn't one "right place" to begin, generally you will suffer fewer complications if you begin the balancing process using a substance that contains the **greatest number** of elements **or** the **largest subscript** values. Noting that you must have at least that many atoms of each element involved, coefficients can be used to increase the "atomic inventory". In the next few questions, you will see one **emboldened** substance in each equation. This emboldened substance is the one that I judge to be a "good" starting place. One last hint--modify the coefficients of uncombined elements, i.e., those not in compounds, <u>after</u> you modify the coefficients for compounds containing those elements -- <u>not before</u>!

3.7. Balanced chemical equations for these reactions:

(a) $2\ Al(s) + Fe_2O_3(s) \rightarrow 2\ Fe(\ell) + Al_2O_3(s)$

1. Note the need for <u>at least</u> 2 Fe and 3 O atoms as products.
2. A coefficient of 2 for Fe (in products) balances the Fe inventory.
3. Al_2O_3 (as product) takes care of the O inventory, and mandates 2 Al atoms.
4. 2 Al would give 2 Al atoms on both sides of the equation.

(b) $H_2O(g) + C(s) \rightarrow CO(g) + H_2(g)$

Note this equation has 2 H atoms both left and right;1 O atom on both sides and 1 C atom on both sides. Hooray—it's already balanced.

(c) $SiCl_4(\ell) + 2\ Mg(s) \rightarrow Si(s) + 2\ MgCl_2(s)$

1. Each $SiCl_4$ (reactant) will require 4 Cl atoms (as product). Since $MgCl_2$ has 2 Cl atoms per unit, we'll need 2 $MgCl_2$ units.
2. A coefficient of 2 for $MgCl_2$ mandates 2 Mg atoms (as reactants).
3. A coefficient of 2 for elemental Mg (in reactants) balances the Mg inventory.

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